LABORATORY EXPERIMENTS

TO ACCOMPANY "MODERN CHEMISTRY"

BY

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PREFACE

THESE laboratory experiments are intended primarily for high schools and others of similar grade. The apparatus described is of a simple character, such as any laboratory may provide.

The greatest problem for the instructor in chemistry is how to obtain from his students earnest and thoughtful work in the laboratory. To assist in securing this result many quantitative experiments have been introduced which will require close attention and accurate work on the part of the student. With the same end in view, the students are required to prepare and preserve many of the commoner salts. This will be found to add materially to the interest of the work, as well as to give clear ideas of general chemical processes and definite combining quantities.

The modern theory of ionization is presented frequently in an experimental way, and so simply that the student cannot fail to grasp the idea. Likewise, mass action, hydrolysis, and the other important laws of chemistry, usually found in similar manuals, are taught by experiment.

In many instances, simple methods have been given for working out experimentally the chemical equations. The purpose of such work is to impress upon the student the facts that these equations are all based upon experiment, and that they are valueless without such foundation.

Furthermore, a considerable amount of work has been introduced for special students. It is always a great question what to do with the stronger pupils who are well able to perform an extra number of experiments. The work intended for these will be found of interest, and will greatly aid the teacher in solving this difficult problem.

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LABORATORY EXPERIMENTS

INTRODUCTION

1. Apparatus. — Every student should supply himself with an apron reaching to the ankles. This may be of denim, oil-cloth, or rubber. He should also have a towel, a bar of soap, a triangular file, a rat-tail file, a small candle, and a box of matches. He will probably be furnished the following articles, which he should keep clean and in order: —

Test-tubes, $5 \times \frac{1}{2}$, $6 \times \frac{1}{2}$, $6 \times \frac{3}{4}$.
2 Evaporating Dishes, medium and small.
2 Beakers.
1 Flask.
1 Test-tube Brush.
1 Pair Forceps.
1 Blowpipe.
1 Glass Rod.
Glass Tubing.
Corks.

Other apparatus, such as pneumatic trough, test-tube rack, Bunsen burner, wire gauze, ring-stand, funnel, etc., may possibly be used in common by a number of students, and kept in a place accessible to all.

2. How to Proceed. — The student should know first the *purpose* of every experiment, and the *steps* he must take to secure the correct results. Do not hurry. Impatience, or attempts to hasten matters, may vitiate the whole work and necessitate a repetition of the experiment. Usually, you will find the results are not given in this book; hence it is very important that you see *everything* which happens.

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Do not draw your conclusions beforehand and attempt to make the results conform to them. If the work is done carefully, as a rule your results may be accepted as trustworthy, regardless of any preconceived ideas.

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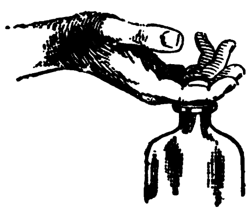

FIG. 1.

Keep the reagent bottles in some particular order, so that you need lose no time in finding them.

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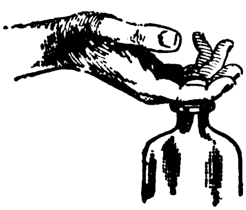

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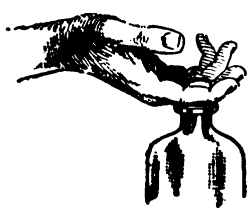

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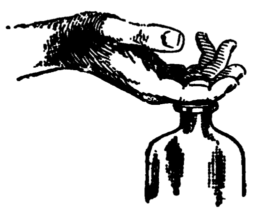

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CHAPTER I

PREPARATORY WORK—GLASS MANIPULATION, ETC.

1. Owing to limited quantity of certain forms of apparatus, all of the students in the class may not be able to work upon the same thing at the same time. The instructor will be governed by the size of the class and the apparatus at his disposal.

2. Cutting Glass Tubes.—With a sharp triangular file make a deep scratch in the glass tubing where you desire to cut it. Seize the tube with both hands, the thumbs below and meeting just under the scratch, the fingers above. Pull strongly apart and bend downward. To remove the rough ends of the tubes, hold them in the Bunsen flame until the glass begins to soften. This must always be done to prevent their cutting rubber connections and corks, and thus forming leaky joints.

3. Larger tubes may be cut in the same way, but the scratch should be made somewhat deeper. Sometimes it may be necessary to lay the file upon the table with the edge just beneath the scratch in the glass, and then press strongly downward with the hand upon either side of the file.

4. Very large tubes or bottles are best cut by unequal heating. One method that is often used is to tie an ordinary twine string around the bottle where it is desired to cut it. Saturate the string with kerosene, and ignite it. When the oil has burned off, pour quickly over the whole bottle a quantity of cold water. The bottle will crack along the line of greatest heat.

5. Bending Glass Tubes.—This is very simple when understood, but cannot be done successfully with the Bunsen burner alone. A broad, flat flame about an inch and a half in width is best. This may be obtained by attaching a wing-top to the Bunsen burner, or by using an ordinary burner with the gas half turned down. (See Fig. 2.) Hold the glass tubing in the flame, rolling it around so as to heat all sides equally. When it is well softened, remove from the flame and make the required bend.

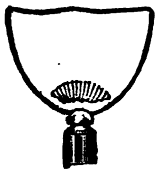

FIG. 2.

When properly done, the turn should be well rounded and not crumpled on the inner side. (See Fig. 3.) Moreover, the diameter should remain the same at all points.

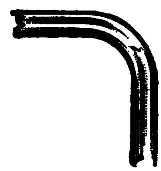

FIG. 3.

6. Before allowing the glass to cool, hold it in the luminous flame of the burner so as to cover it well with soot. This will necessitate a slower cooling, and render the glass less liable to crack.

7. Glass Jets.—For testing the combustibility of gases, a tube drawn to a fine point is desirable. Take a piece of tubing 6 or 8 inches in length and hold it in the flat flame near the middle of the tube. Roll it around, and when well softened remove it from the flame and pull it slowly apart, until you have a tube of small diameter where heated. (See *a* in Fig. 4.) When cooled somewhat, cut in two at *a*, round off the edges as previously described, and anneal. In one of the tubes

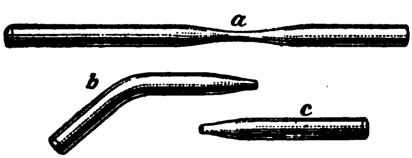

FIG. 4.

make a bend about three-fourths of an inch from the larger end, as shown at *b*.

8. Specimen Tubes. — The student may desire at times to preserve certain chemicals that he prepares. For this purpose, specimen tubes which he can make himself serve excellently. For solids, these tubes may be 3 to 4 inches in length, and for the sake of appearance should be uniform in size. Cut off a piece of glass tubing twice the length desired for the completed tube, and round off the sharp edges.

9. When cool, the glass should be thoroughly cleansed inside. Put a cork into one end of the tube, and drop in a little powdered potassium dichromate, then add a little sulphuric acid. Now cork the other end and tip the contents back and forth a few times, until all parts have been well moistened. Rinse out the tube and dry thoroughly.

10. When the clean tubes are dry, with a moderately small flame from the blast lamp, heat the tube at the center, constantly rolling it around. When the walls have collapsed from their own weight, as shown at *a* in the figure, remove from the flame and suddenly draw it apart. You will now have two tubes, each ending in a long capillary tube, as *b* and *c*. Turn down the gas to a very small flame and direct it against the shoulder of one of the tubes, and, when soft, again draw out the capillary tube. With two or three successive heatings

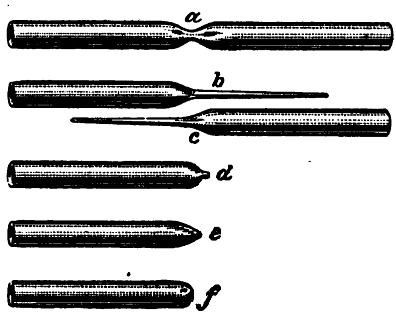

FIG. 5.

in this way you will have a tube with very little surplus glass at *c*. Then fuse off the capillary at this point.

11. The tube will now have a little knob, as at *d*. Turn on the gas more strongly, and heat the tip at this point till red hot. Remove from the flame and blow in the end of the tube until the glass is of the same thickness everywhere; then heat the whole of the bottom, *e*, and again blow, rounding out the bottom. Anneal as usual.

12. To seal Tubes. — Many liquids and some solids should be sealed hermetically for preservation. To do this, longer tubes than those previously described are necessary. Heat cautiously the open end of the tube until the glass softens, then attach a warm glass rod to be used as a handle. (See *a* in the figure.) Turn down the gas to a fine jet, and heat a narrow ring, letting the tube collapse as previously described. When well softened, draw out the tube a short distance, as shown at *c*, and hold steady until the glass has become stiff; then quickly fuse in two at *d*, closing the tube as shown at *e*. Anneal as usual. If it is desired to seal volatile or inflammable liquids, the tube must be considerably longer and wrapped in a mixture of ice and salt. The operation may then be conducted with safety.

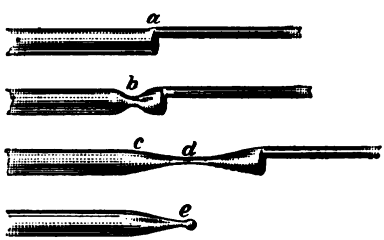

FIG. 6.

13. Repairing Test-tubes. — Broken test-tubes may be easily repaired by the student. The plan is practically the same as that used in preparing and sealing specimen

tubes. As the glass is very thin and with ordinary tubes soft, but little heat is necessary.

14. Opening Bottles. — Many of the common reagents are received in glass-stoppered bottles, sealed with wax or paraffin. This may be melted by holding the neck of the bottle over a low gas flame, and rapidly rolling the bottle in the hands so as to heat the glass slowly and evenly. The wax is soon softened and the stopper easily removed. With a little care there is no danger of ever breaking a bottle. Always be sure, however, to avert the face in removing a stopper after heating the glass in this way.

15. Filtering. — There are two methods of folding filter paper; the first is always used when it is desired to remove the precipitate from the paper. Fold the paper to a semicircle, and then to a quadrant, making one fold a little larger than the other. Open out the larger "quarter," for this usually fits better than the smaller, and put into the funnel. You will thus have three thicknesses on one side and one on the other. Moisten the paper a little and press snugly against the funnel so as to remove any air bubbles. A more rapid filtration is secured by folding the paper back and forth like a fan, so as to expose a large amount of surface to the filtering liquid.

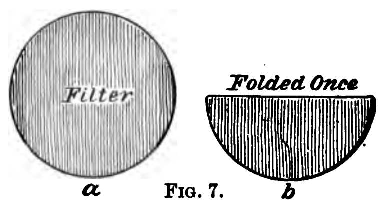

FIG. 7.

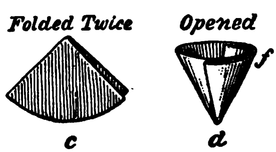

FIG. 8.

16. The next step is the proper handling of the liquid itself. As shown in the figure, let the stem of the funnel

touch the side of the receiving beaker. Then moisten a glass rod and slowly pour the liquid down this upon the filter. For more rapid work, see "Modern Chemistry," page 370.

Fig. 9.

17. Decanting. — If the precipitate is one which readily settles to the bottom, and small quantities of the liquid in it will not interfere with subsequent work, the solution may be carefully poured off or *decanted* from the precipitate. The advantage of this is its rapidity.

18. Work with the Blowpipe. — Examine your Bunsen burner, light it, and notice the effect of opening and closing the holes at the bottom. Turn down the gas to a small flame, and close the openings at the bottom. (See *a* in the figure.) For the *reducing* flame put the tip of the

Fig. 10.

blowpipe just within the outer edge of the flame, and blow gently. Practice until you can breathe and blow at the same time. There should be a small luminous cone in the center of the horizontal flame (see *b*). It is the red-

hot carbon in this portion of the flame that effects the reduction of metallic oxides. (See "Modern Chemistry," page 139.) Regulate the size of the flame and the amount of air forced into it until you secure this luminous cone.

19. For oxidizing work, put the tip of the blowpipe into the center of the flame, as shown by *c* in the figure. Regulate the flame and your breath so that now you secure a perfectly non-luminous cone. As the reducing flame contained an excess of hot carbon, so this has an excess of oxygen, which it readily imparts to any oxidizable substance. It is, therefore, called the *oxidizing* flame.

20. For practice, the student may put a small amount of lead oxide (litharge) into a cavity made near one end of a stick of charcoal and heat with the reducing flame till the metal is obtained; then heat with the oxidizing flame until the litharge is obtained again.

21. Collecting Gases. — The method of collecting gases depends upon the character of the gas and the purpose for which it is to be used. Such gases as are insoluble in water are generally collected *over water*. A little water is put into the pan which serves as a pneumatic trough, and the bottles, filled with water, are inverted in the pan. To do this, hold a sheet of glass or paper over the mouth of the bottle, and invert quickly. When the bottle has been lowered into the water, remove the sheet of glass.

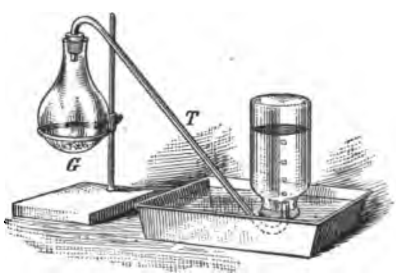

Fig. 11.

22. Downward Displacement. — Gases which are soluble in water cannot be collected as described above. If the

use to which they are to be applied require them to be unmixed with air, they may be collected over mercury, provided they do not combine with this element. Usually, however, such precautions are unnecessary, and heavy gases are collected by *downward displacement.* The gas, as shown by *a* in the figure, is introduced below the air, and gradually forces the latter out of the bottle.

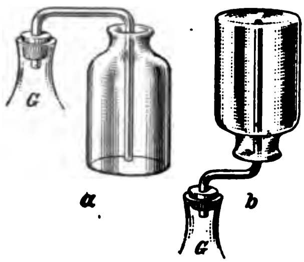

FIG. 12.

23. Light gases, soluble in water, are collected by *upward displacement,* as shown by *b* in the figure.

24. Volumetric Measurements. — The student will have occasion constantly to use the cubic centimeter as his unit of volumetric measurement. He should, therefore, have a fairly accurate idea of this unit. This can come only by practice.

25. Fasten a burette upon a stand and fill it with distilled water; see that the tube below the stop-cock is completely filled. Allow 1 cc. of water to flow into a 6-inch test-tube and fasten a narrow strip of mucilage paper even with the lower edge of the meniscus. Run in another cubic centimeter, and mark it in the same way. Thus mark each up to five; indicate also 10 cc. (See Fig. 13.) Now pour out the water, and with a sharp file make a neat scratch just above each strip of paper. Now remove the paper. Make a drawing of the graduated tube, full size,

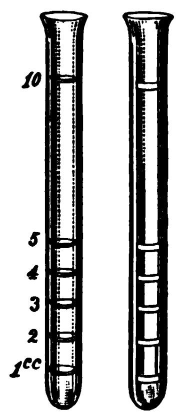

FIG. 13.

26. Gravimetric Work. — Put a beaker upon the left scale pan of the balance, and counterpoise it with a box of shot or sand upon the other. Add a 5-gram weight to the shot, and from a burette allow water to run into the beaker until the weight is balanced; mark the height of the water as in graduating the test-tube. Proceed thus for 10, 15, 20, and 25 cc., and make the work permanent with a file as before.

27. The Water Bath. — Most laboratories will not be able to furnish a regular water bath to the students. But unless the work is to be of long duration, a small saucepan, or large cup, of granite ware, answers equally well for much of the work. With a few covers, made from galvanized iron, and having openings varying in size, such a water bath will serve nearly all purposes.

28. Platinum Wires. — These are used for *flame* and *bead tests*. Usually, for convenience, the wire is fused into a short glass tube. Prepare a jet tube, as already described, and insert for a short distance the end of a platinum wire. Fuse the glass in the Bunsen flame, and hold in position till the tube has slightly cooled.

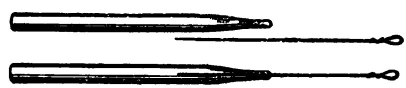

FIG. 14.

CHAPTER II

CHEMICAL CHANGES — MIXTURES

1. Chemical Changes. — What is a physical change? A chemical change? Put into a hard glass tube about 1 g. of mercuric oxide, HgO, and heat strongly for some time. Note any changes in appearance. What deposit forms upon the sides of the tube? Put a pine splinter with a spark upon it into the tube; result? What kind of a change is going on? Why do you say this?

2. Reaction of Sulphur with Potassium Chlorate. — (Class room.) What have you when you first put together a pinch of sulphur and powdered potassium chlorate? What happens when they are rubbed vigorously? What have you at the close? What kind of a change? Why?

3. Reaction of Phosphorus with Potassium Chlorate. — (Class room.) Fold up in a piece of paper a minute piece of phosphorus and a pinch of potassium chlorate. State effects of striking with a weight.

4. Sugar, Potassium Chlorate, and Sulphuric Acid. — (Class room.) Put about a gram each of sugar and of finely powdered potassium chlorate into an iron saucer. Stir them together; what have you now? Why? Drop upon the mass, by means of a pipette, a little strong sulphuric acid; result? What kind of a change has taken place?

5. Put into a porcelain mortar a few crystals of potassium iodide and about the same amount of mercuric chloride. Rub them strongly together; result? Explain. Add water, and after standing a few minutes pour into a

beaker. Set aside for further study. Put 2 or 3 g. of sugar into an evaporating dish and just moisten with water. Warm gently, and add cautiously a little strong sulphuric acid; result? Explain.

6. **Mixtures.** — What is the difference between a mixture and a compound? In section 5 above what did you have before rubbing together? Afterward? Are there still two substances? Determine thus: pour off the clear solution into an evaporating dish and boil nearly to dryness. Does the liquid contain any substance in solution? Describe.

7. **Separating a Mixture by Crystallization.** — Sometimes a mixture may be separated by boiling down the solution of the two substances, until one will crystallize out. Put into a beaker 6 g. of potassium chloride and 6 g. of potassium chlorate. Add 100 cc. of water, and warm to dissolve. Now, upon a sand bath, boil down slowly till you have only about 50 cc. of the solution, and allow it to cool.

8. While waiting, put a small crystal of each of the compounds used above into separate test-tubes, and dissolve in about 1 cc. of distilled water. To each add a few drops of silver nitrate solution; result?

9. When the solution has cooled, filter out the crystals. (See instructions, pages 11, 12.) Wash several times with cold water. Put one or two into a test-tube, dissolve in water, and test with silver nitrate solution; what are the crystals? Test, in the same way, the solution filtered from the crystals; what is it?

10. **Separation of Mixtures by Solution.** — How have you already applied this method? Try also gunpowder. Put into an evaporating dish 2 or 3 g. of gunpowder and add 10 or 15 cc. of water. Warm gently for a few minutes.

Filter, and in an evaporating dish boil the filtrate nearly to dryness.

11. While waiting for the above, transfer the black solid from the filter paper to an evaporating dish and add about 5 cc. of carbon disulphide, CS_2. Shake two or three minutes and decant the clear solution into an evaporating dish, and let it dry in the hood without heating. Carbon disulphide is very inflammable. It may be necessary to rinse the black residue a second time with carbon disulphide. What is the substance not dissolved by the last reagent?

12. Describe the solid left from each solution above. How many ingredients of gunpowder have you found?

13. If you had a mixture of sulphur and iron filings, how could you separate them? Two methods? Why?

14. Separation by Distillation. — See work in Carbon, page 50.

15. Compound of Iron and Sulphur. — Mix together 7 g. of fine iron filings and 4 g. of flowers of sulphur. Put them into a hard glass tube and heat strongly, better in a blast lamp, until a red glow passes through the mass. Cool and break open the tube. Determine whether you can separate the filings and sulphur by either of the methods suggested above. Give reasons.

CHAPTER III

WATER

1. **Deliquescent Substances.** — Put into a dry beaker or evaporating dish a small piece of fused calcium chloride; into another, some potassium hydroxide. Allow them to remain exposed to the air during the time you are at work in the laboratory. Before leaving, notice whether any change has taken place. Allow them to remain thus over night, and note changes. Can you explain? Such substances are said to be *deliquescent*. Can you suggest any use for them?

2. **Efflorescent Substances — Water of Crystallization.** — Examine two specimens of ferrous sulphate; the one from a bottle recently opened, the other, one which has been exposed for some time to the air. Note two differences in appearance. Substances which change thus when exposed to the air are called *efflorescent*. Let us determine the cause of the change.

3. Put into a test-tube two or three small crystals of ferrous sulphate, and heat gently for a few minutes. What two changes do you see taking place? What do you see collecting upon the cooler portions of the tube?

4. Put into another dry tube a few very small crystals of copper sulphate, and heat gently as before. Note two changes. What collects upon the tube? After the test-tube has cooled, add two or three drops of water to the contents of the tube; results? Explain loss of color and restoration.

5. What is water of crystallization? How can you remove it?

6. Amount of Water of Crystallization. — Determine the weight of a small porcelain crucible or evaporating dish. Weigh out into it 1 g. of magnesium sulphate, and support the crucible in a triangle upon a ring-stand. Heat gently at first, then to dull redness. Cool and weigh. Repeat the operation until two successive weighings give the same result. The loss indicates the water of crystallization expelled by the heat. What per cent of magnesium sulphate, as shown by your results, is water of crystallization? Repeat the experiment to verify your conclusions. Tabulate the results.

7. If thought desirable, the instructor may furnish other suitable crystallized compounds for similar work. Copper sulphate and alum serve well. Crystals should be chosen which have not effloresced. They should be powdered and weighed quickly.

8. Composition of Water. — Proof of the quantitative composition of water will usually be made by the instructor in the class room. Let the student make drawings of the apparatus used both for the analytic and synthetic proof, and state exactly what was done. If suitable apparatus is not at hand, the students may prepare it themselves. (See "Modern Chemistry," pages 367, 368, and 32.)

FIG. 15.

9. By means of a simple electrolytic apparatus, which the students can easily make (see "Modern Chemistry," page 368), collect a test-tube full of the mixed gases, and bring a lighted match to the tube; results? You can thus readily

determine whether you have the two gases separate or mixed.

10. Solubility of Salts in Water. — (For special students only.) To plot the solubility curve of potassium alum or ammonium chloride. Measure into a large beaker or flask 100 cc. of distilled water, put it upon a sand bath, and gradually heat till it begins to boil. Meanwhile add slowly, with constant stirring, finely powdered alum as long as any will dissolve. When saturated add another small lump. Weigh a small evaporating dish, and by means of a pipette transfer 10 cc. of the saturated solution, at 100° C., to the dish, and evaporate on the water bath to dryness. When dry, weigh and determine the amount of alum contained in the 10 cc. of water at 100° C.

11. While this has been evaporating, you should allow the remainder in the large beaker to cool. When it reaches 90° C., again measure out 10 cc. of the solution into another weighed dish, and evaporate to dryness as before. Determine the amount of alum in this 10 cc.

12. Proceed thus, cooling the solution to zero, or as low as you can, estimating the amount of alum in 10 cc., at 80°, 70°, etc. Next, determine from these data the amount of alum that would be dissolved by 100 cc. of water at these different temperatures.

13. To plot the Solubility Curve. — Draw on the opposite page, near the left edge, a vertical line 4 inches long, and a horizontal one from its lower end the same length. See Fig. 16, which illustrates the method with an imaginary compound. Divide each of these lines into ten equal parts, and draw horizontal and perpendicular lines. Let the vertical spaces represent the temperature, and the horizontal the grams of salt dissolved. Next, put a small dot upon each horizontal line at a point which would

be intersected by a vertical drawn from a point on the base line representing the number of grams. For example, suppose at 0° you find 100 cc. of the solution contains 3 g. of the salt, put a dot on the base line three-fifths of the distance from 0° to 5 g.; if at 10° you find 100 cc. of the solution contains 5 g., put a dot at the intersection of the 10° and the 5 g. lines, as in the figure.

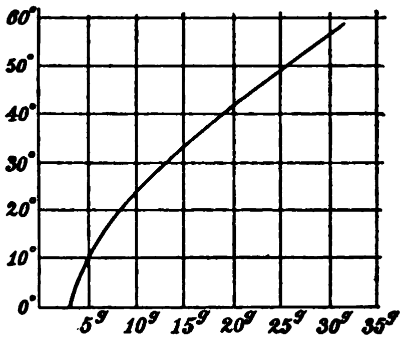

FIG. 16.

Proceed thus for all the observations made, and connect the dots with a curved line.

CHAPTER IV

HYDROGEN

1. **Preparation from Water.** — How have you seen the hydrogen obtained from water? What would have been the results had you used pure water? If you had passed the current through concentrated sulphuric acid? Endeavor to find some explanation of these facts, if possible. Look up the theory of "electrolytic dissociation" or ionization.

2. Hydrogen may also be obtained from water by using metallic sodium. Unless the classes are small, the instructor should do this work, as the handling of sodium is somewhat dangerous.

3. Put a very small piece of sodium, not larger than a grain or two of wheat, upon a vessel of water. Note the changes. Touch a light to it as it rolls around; results? What does the flame prove?

4. Put a piece of sodium not larger than a pea into a gauze spoon, and thrust under a test-tube of water, as shown in Fig. 17. When the tube is filled with gas, test it as you did when hydrogen was obtained by electrolysis; results? Does it seem to be the same gas? Have you any evidence that the oxygen has also been evolved from the water as before? Is there any evidence to the contrary? If so, what? (See page 20.)

FIG. 17.

5. The Chemical Action. — Examine the water over which the hydrogen was collected. First, put into a test-tube about 1 cc. of sodium hydroxide solution and add a drop or two of an alcoholic solution of phenol phthalein. Try a test-tube of pure water with a drop of the phenol. Finally, put a drop of the phenol into the water from which the hydrogen was obtained. State your conclusions in the matter. Write an equation, expressing the above facts: —

Water + Sodium = + .

6. Hydrogen from Acids. — The above method is too slow when considerable quantities of gas are required.

7. Fit up a flask, delivery tube, and pneumatic trough, as shown in Fig. 18. The flask need not have a capacity of more than 100 cc. or 150 cc. Put into it about 8 g. of granulated zinc, and nearly cover with water. Now add slowly a little concentrated sulphuric acid, and, if you are not using a cork with thistle tube,

quickly close the flask. When the gas has been escaping rapidly for two or three moments, collect a test-tube full of it. (See instructions on page 13.) Test this as you have before. If it burns quietly, fill three bottles, about 200 cc. each. Do not throw out the contents of the flask.

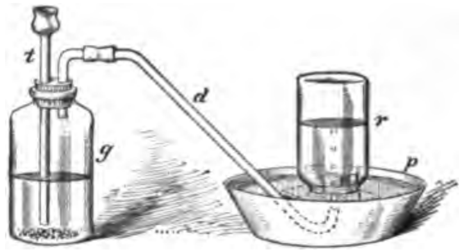

FIG. 18. — Hydrogen Apparatus.
p = pan. g = generating flask.
t = thistle tube. r = receiving flask.
d = delivery tube.

8. Holding the first bottle inverted, but slightly inclined, thrust up into it a lighted candle. Notice the effect of the hydrogen upon the flame. Withdraw the candle from the bottle; results? Why? Repeat. State four facts regarding hydrogen learned by this experiment.

9. When all the hydrogen is out of the first bottle, put the lighted candle up into it, and notice that it continues to burn as in the air. Take a second bottle of hydrogen and gradually pour its contents into the first; how will you do this? Test the bottle into which you have poured the gas; does it contain hydrogen? Test the other bottle; results? What does this experiment show?

10. The third bottle may be used with the hydrogen pistol, if the class is very small. (See "Modern Chemistry" for details.) The experiment may be varied, as follows: Fill a test-tube two-thirds full of water and invert it over the pneumatic trough. Now pass in hydrogen until the tube is full; to the mixture of gases bring a lighted match; results? conclusions?

11. Replace the delivery tube in the hydrogen generator, used for the preceding work, by a jet. (See instructions, page 8.) If necessary, add a little more acid to renew the evolution of gas. Allow the hydrogen to escape a moment or two, then by means of a delivery tube connected temporarily with the jet collect a test-tube of the gas. If it burns quietly, remove the delivery tube and light the jet.

CAUTION. — Never touch a light to a jet of hydrogen before testing a sample in a small tube. Why?

12. What color has the burning jet of hydrogen? Did it have that color at first? If you did not notice, blow out the flame and relight it. When before have you seen burning hydrogen have the same color? Explain.

13. Hold a cold dry beaker or bottle over the flame; what collects? Explain its formation. Hold bottles of various sizes over the flame, slowly moving them up and down; results? Study the flame carefully, meanwhile; can you explain the "singing tones"?

14. Name the characteristics of hydrogen you have learned by the preceding experimental study of it.

15. Chemical Changes Involved. — If any pieces of zinc remain in the generating flask, remove them, and after washing them, preserve for future use. Filter the solution to remove any sediment, and evaporate the clear filtrate to less than half the original quantity. Set it aside for twenty-four hours; semi-transparent crystals should form. Filter them out, wash with a very little cold water, and again concentrate the filtrate and wash water till further crystals will form upon cooling. After washing these as before, add them to the others and, when thoroughly dry, preserve in a specimen tube.

16. This compound is zinc sulphate, $ZnSO_4$. Determine whether it contains water of crystallization; results? (See page 19.) Can we prove that the compound is zinc sulphate? Dissolve a few crystals in 5 or 6 cc. of water; according to the theory of electrolytic dissociation (ionization) the molecules are now, to a considerable extent, broken up into zinc ions and sulphate ions, and may be detected by adding certain reagents.

17. Test for the Zinc Ions. — To about half of the solution prepared above, add slowly, drop by drop, a solution of potassium hydroxide. If zinc ions are present, a gelatinous, white precipitate will form. Conclusions?

18. Test for Sulphate Ions. — First add a drop of sulphuric acid to 3 or 4 cc. of water. You have probably learned (page 22) that while concentrated sulphuric acid is not ionized to any great extent (known by the fact that it conducts but little electricity), quite the contrary is the case with dilute acid. We now have a greatly ionized solution; to it add a little barium chloride solution; results? To this, add a little nitric acid; results? In the same way, test for sulphate ions in the solution of crystals prepared above. Conclusions?

19. Statement of Results. — By putting together zinc and sulphuric acid, you have obtained what two products? Express these facts briefly, thus: —

$$Zn + H_2SO_4 = \quad + \quad .$$

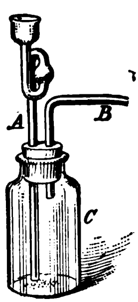

FIG. 19.

20. Displacing Power of Metals. — You have seen that zinc has the power of displacing the hydrogen in sulphuric acid. What is true of zinc and sulphuric acid may be shown also for various other metals with other acids. Arrange a generating flask as shown in Fig. 19, except

that you may use a straight tube instead of the one shown at *A*. Over a rather deep pneumatic trough, invert a 2-liter bottle filled with water, and arrange to pass the hydrogen into the bottle.

21. Put into the generating flask 1 g. of aluminum wire and pour in strong hydrochloric acid slowly, so as to force no air into the bottle. Action will soon begin and become vigorous. When all the metal has been dissolved and the generating flask has cooled, raise or lower the bottle of hydrogen until the water stands at the same level within and without; slip a sheet of glass under the mouth and place in an upright position. Ascertain the volume of hydrogen set free by the aluminum, by measuring the water required to fill the bottle. This gives the volume of the gas at the temperature of the room. For method of reducing to standard conditions, see "Modern Chemistry," pages 96 and 374. Suppose the volume as measured is *n* cc. and as reduced is *m* cc. It has been found that 1000 cc. of hydrogen weigh .0896 g. Find the weight of *m* cc.

$$\frac{m \times .0896}{1000} = w,$$

grams of hydrogen displaced by 1 g. of aluminum. Then

$$\text{wt. of H} : \text{wt. of Al} : : x : 27 \text{ (combining wt. of Al)},$$

$$x = ?$$

22. Let a portion of the class try the same experiment, using magnesium instead of aluminum, another portion, zinc, and a fourth, iron. Find the amount of hydrogen displaced by each of these metals, using as their combining weights, aluminum = 27, magnesium = 24, zinc = 65, and iron = 56. (See "Modern Chemistry," p. 9.)

CHAPTER V

OXYGEN

1. To prepare the Gas. — Arrange apparatus as shown in Fig. 20, and prepare to collect the gas over water. (See directions, page 13.) A 6-inch test-tube may be used instead of the flask. Counterpoise a small filter paper upon the scales and add to it about 1 g. of manganese dioxide. Weigh accurately. Add to this 2 or 3 g. of potassium chlorate, mix the two compounds, and transfer them without loss to the generating flask. Save the filter paper. Make the connections tight, and heat gently. Allow the first gas that comes over to escape. Continue to heat moderately so as to secure an evolution of gas, not too rapid, and fill two 500 cc. bottles and two or three smaller ones. When no more gas can be driven out even by strong heat, disconnect the delivery tube and *then* remove the flame. What is the danger in not doing this? When the flask has cooled, add to it 30 to 50 cc. of water, cork it and set it aside. With the gas collected make the following experiments. After each, add a little water to the bottle used, mark it and set it aside.

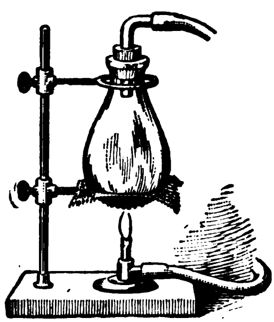

FIG. 20.

2. Test one of the small bottles with a glowing pine splinter, or a candle; results?

3. Heat a stout platinum wire with a loop in the end in the Bunsen flame and dip it into some sulphur. Ignite the sulphur and quickly thrust the wire into one of the smaller bottles of oxygen; results?

4. Fasten a piece of bark or sawdust charcoal upon an iron wire, heat it red hot and quickly insert in the third small bottle; results?

5. Warm the end of a watchspring or iron wire and dip it into sulphur. Ignite it and quickly thrust into a large jar of oxygen, which should have about an inch of water in the bottom; why? Results?

CAUTION. — Phosphorus is dangerous; be careful!

6. Put into a deflagrating spoon a piece of phosphorus about half as large as a pea, ignite it, and insert in the other large bottle of oxygen; results?

7. To test the Products Formed. — To a piece of blue litmus paper touch a drop of the solution in each of the bottles you have set aside; results in each case?

8. Take a fresh piece of blue litmus paper, and learn the effects upon it of sulphuric, hydrochloric, and any other acids you may have on your table; results? What was the nature of the solutions in the bottles? What is the meaning of the term *oxygen?* Why so given? What kind of compound does it form when it combines with another substance? What kind of *oxides* are most of those obtained above?

9. Determination of the Chemical Equation. — It is customary to express the chemical changes that take place, when two or more substances react upon each other, by an equation. The first member gives the substance used; the second, the products obtained; and they are all determined by experiment, as you have done in one or two preceding cases. (See page 26.) In preparing

oxygen you used potassium chlorate and manganese dioxide, and obtained oxygen and a dark-colored residue. unknown by you. We may express this as follows: –

$$KClO_3 + MnO_2 = O + x.$$

Can we determine what x contains?

10. The water added to the generating flask (section 1) will have loosened the contents so that now they may be easily removed. Put the filter paper used in weighing above into a funnel (see page 28), and pour upon it the liquid in the flask. Carefully rinse out upon the paper all adhering particles, and wash the black residue twice with a little pure water. Dry the black solid thoroughly and weigh again. How does the weight now compare with that indicated at first? Save this residue for further tests.

11. Put the filtrate and washings obtained above from the black residue into an evaporating dish and boil slowly to dryness. Describe the product obtained. We see then that the black residue contained at least two substances. Let us substitute $y + z$ for x in the equation —

$$KClO_3 + MnO_2 = O + y + z.$$

12. What are y and z? Remove the solid from the evaporating dish, mix it with about half as much manganese dioxide, and heat in a test-tube. Test carefully with a pine splinter to learn whether any oxygen is being evolved; results? May we conclude that the first heating removed all the oxygen, and that now we have potassium chloride, KCl, and not potassium chlorate, $KClO_3$? We shall soon prove this by finding the weight of the oxygen evolved. We may say, then, that $y = KCl$.

13. Take half the black residue on the filter paper and mix with about an equal portion of potassium chlorate.

Heat in a small test-tube ; is oxygen obtained ? What do you conclude as to the black compound ? Complete the equation by substituting values of the unknown quantities, as found by experiment.

14. Weight of a Liter of Oxygen. — (For special students.) Arrange the apparatus as shown in Fig. 21. For *T*, a small flask may be substituted. The bottles, *A* and *B*, should hold 1 l. or more. Weigh the flask and then put into it about 2.5 g. of potassium chlorate, weigh carefully ; add about 1 g. of manganese dioxide and find total weight. Tube *c* should have rubber connections and pinch-clamp at this point, and should extend nearly to bottom of both *A* and *B*. Fill *A* nearly full of water, *c*, completely. Fasten clamp. Fill *B* to same height as *A*, open the clamp and empty *B*.

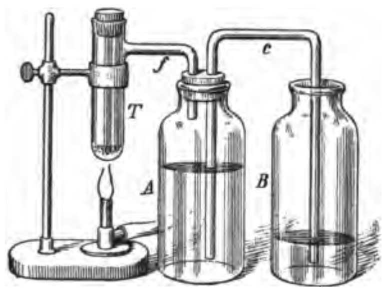

Fig. 21.

15. Heat the potassium chlorate moderately so long as any water is forced from *A* into *B*, allow apparatus to cool, and weigh the flask. The loss represents the oxygen expelled. What ratio does it bear to the potassium chlorate used? Assuming the combining weight of oxygen to be 16, how many atoms of oxygen are represented by x? Does this verify conclusions in section 12?

16. Equalize the level of the water in the bottles as on p. 27, and measure that in *B* to learn the volume of the oxygen at the temperature and pressure of the room. Reduce to standard conditions as you did in case of hydrogen. Suppose the loss in weight of the flask was m g., and the corrected volume of the oxygen, n cc. Find the weight of one 1 l.

CHAPTER VI

NITROGEN AND COMPOUNDS

1. **Preparation.** — What is the usual method of obtaining nitrogen from the air? What objection would there be to using a candle for this purpose? Describe the class-room experiment.

2. Put about 2 g. each of sodium nitrite, $NaNO_2$, and ammonium chloride, NH_4Cl, into a small flask, 100 cc., and add about 25 cc. of water. Start the action by a little heat, and collect a bottle of the gas over water. Keeping the bottle inverted, thrust a burning candle up into it; results? Name as many as you can of the characteristics of nitrogen learned from this experiment.

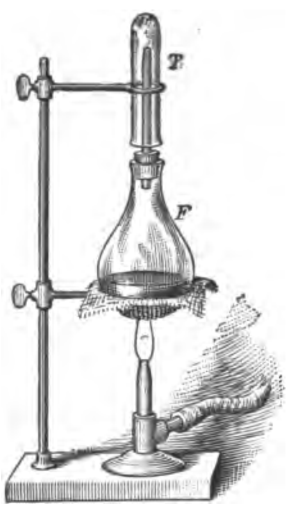

FIG. 22.

AMMONIA

3. **Study of its Characteristics.** — The easiest method of obtaining ammonia is by heating its solution in water. Arrange apparatus as shown in Fig. 22. *F* is a 100 cc. flask, supported upon an iron ring-stand, and protected by wire gauze or a sand bath. *F* is fitted with a cork containing a straight, vertical delivery tube, about 6 inches long. *T* is a test-tube supported over the flask.

4. Put into *F* about 25 cc. of strong aqua ammonia, insert the cork with delivery tube, and put *T* into position. What do you call this method of collecting gases? Warm the flask gently; what does the liquid soon seem to do? Feel of the flask. Explain. Hold a piece of moistened red litmus paper near the mouth of the tube; results. When well filled this change will take place quickly. Thrust a small candle up into the gas; results? Give five characteristics of ammonia learned by your work thus far.

5. Fill another bottle with the gas and invert it over a pan of cold water; after a few minutes note anything that has happened. Explain.

6. **Preparation of Ammonium Hydroxide.** — What is the source of our commercial supply? To illustrate its preparation, support a small flask upon a ring-stand as in Fig. 22. Insert a cork with a bent delivery tube, which reaches down into a bottle partly filled with water. Let the delivery tube just dip beneath the surface of the water; why? Put into the flask a mixture of about 5 g. of lime, CaO, and 10 g. of ammonium chloride, NH_4Cl, and add about 10 cc. of water. Warm gently for some time. Notice the odor of the water in the bottle after a few minutes. What is its effect upon litmus paper? What is the solution?

7. **Neutralization. Compounds of Ammonia.** — Refer to work in oxygen; what have you found is the effect of acids upon litmus paper? What is the effect of ammonia?

8. **Ammonium Chloride, NH_4Cl. Preparation.** — Put into a beaker 25 cc. of strong hydrochloric acid and add to it very slowly with constant stirring ammonium hydroxide until a drop of the solution will not affect either red or blue litmus. The acid and the ammonia have *neutral-*

ized each other, so that neither possesses its former characteristics.

9. If the solution has any sediment, allow it to settle, then decant or filter. Evaporate slowly on a water or sand bath till you have about 15 cc. of the liquid. Cool, and drain off the liquid from the crystals which form; evaporate this mother liquor still further and add these second crystals to the first. If not pure white, dissolve in a *very little* boiling water and allow to crystallize again. Drain off, dry thoroughly, and preserve in specimen tube. Dissolve one or two small crystals in a few drops of water and test it with litmus paper; results? Put two or three small crystals into a dry test-tube and heat gently; results? Do you obtain any evidence of water of crystallization? Describe the compound prepared above and state what characteristics you have learned.

10. Ammonium Nitrate, NH_4NO_3. — (Special work.) Put 25 cc. strong nitric acid into a beaker and surround with ice or snow. Add ammonium hydroxide very slowly with constant stirring and avoid letting the solution rise much in temperature. When a drop of the liquid begins to turn litmus paper blue, set the beaker aside over night. Decant from any sediment and evaporate upon a sand bath till a drop of the liquid upon a stirring rod will upon cooling deposit crystals. Allow it to cool. Drain off the mother liquor from the crystals and evaporate further. Put all the crystals together, dissolve in very little hot water, and recrystallize. Dry and preserve. Describe the compound.

Oxides of Nitrogen

11. Nitrous Oxide, Nitrogen Monoxide, N_2O. Preparation. — Put into a test-tube about 2 g. of ammonium

nitrate, NH_4NO_3, and arrange to collect the gas over warm water. Why? Clamp the test-tube (Fig. 21) upon a ring-stand and heat gently — just enough to secure a slow evolution of gas. Fill two small bottles and test as you did oxygen, with a glowing splinter, a burning candle, and with phosphorus; results? What is the odor of the gas? Compare, in tabular form, with oxygen.

12. Chemical Action. — (Special students.) Devise some plan, and arrange apparatus to prove by experiment that when ammonium nitrate is heated both water and nitrous oxide are formed. Can you also prove quantitatively the equation —

$$NH_4NO_3 = N_2O + 2\,H_2O\,?$$

Use about 1.6 g. ammonium nitrate. It will not be necessary to decompose the entire amount used.

13. Nitric Oxide, Nitrogen Dioxide, NO. Preparation. — Put into a small flask a few grams of copper turnings and add moderately dilute nitric acid, till the copper is nearly covered. No heat is necessary. Collect two or three bottles of the gas evolved over water. Save the contents of the flask. What do you notice in the flask at first? Has the gas any color? Open a bottle of it; results? Can you explain what you saw in the flask at first? What you see is nitrogen peroxide, or tetroxide, NO_2, formed by the union of atmospheric oxygen with the nitric oxide, thus: —

$$NO + O = NO_2.$$

14. Quickly invert over the pan of water the bottle filled with the colored fumes; results? Why? Explain why the fumes which appeared in the flask at first soon disappeared.

15. Into another bottle of the nitric oxide lower a burning match or candle; results? Try a third bottle with phosphorus, burning strongly in a deflagrating spoon; results? How do you account for this in view of previous conclusions?

16. Name some respects in which nitric oxide is like oxygen; one in which it is different from any other gas studied.

17. The Chemical Changes. — We shall not attempt to work out the above reaction in full. Decant the clear blue solution in the flask from any residue, or filter, if necessary. Evaporate it slowly on a water or sand bath to small volume, and allow it to crystallize. Usually, if the solution is sufficiently concentrated, it will crystallize as a solid mass. Break up somewhat and dry in a desiccator over potassium hydroxide. This is cupric nitrate. Preserve in a specimen tube.

18. Two of the products obtained, therefore, in the above work, are nitric oxide, NO, and cupric nitrate, $Cu(NO_3)_2$. Heretofore you have found that when an acid and a metal react upon each other, hydrogen was produced, and you would look for the same here. Probably it is formed, and may be expressed thus: —

$$3\,Cu + 6\,HNO_3 = 3\,H_2 + 3\,Cu(NO_3)_2;$$

but the nascent hydrogen decomposes two additional parts of nitric acid, thus: —

$$3\,H_2 + 2\,HNO_3 = 2\,NO + 4\,H_2O.$$

Adding the two equations, we have, —

$$3\,Cu + 8\,HNO_3 = 2\,NO + 3\,Cu(NO_3)_2 + 4\,H_2O.$$

19. Nitric Acid. Preparation. — This is better done in the class room unless the classes are small. Into a glass-stoppered retort put a handful of sodium or potassium nitrate, and cover with strong sulphuric acid. Avoid all rubber connections. Let the neck of the retort extend into a flask kept cool by ice, or by a wet cloth or paper over it. Shake the contents of the retort, so as to mix them, and distill slowly on a sand bath for several hours. Examine the product obtained, and describe it.

20. Behavior of Nitrates. — Powder about a gram of the cupric nitrate obtained in preparing the nitric oxide, and put it into a test-tube. Attach a delivery tube, and arrange to collect the gas in a small bottle or a test-tube over water. Heat moderately until the blue solid is entirely converted into a black powder; what gas did you see evolved? Test the gas in the bottle with a glowing pine splinter; results? What is the black residue? Write the reaction.

21. In the same way try a little lead nitrate; results? Reaction? Try also sodium nitrate in the same way; how are the results different? You would find that potassium nitrate gave the same results. How have you found ammonium nitrate to act when heated? (See page 35.) What can you say of the stability of nitrates as a whole?

22. Test for Nitric Acid or a Nitrate. — Dissolve the substance, if a solid, in a few cubic centimeters of water, and drop into it a small crystal of ferrous sulphate. Then pour down the side of the tube cautiously, so as not to mix the acid with the solution, a little strong sulphuric acid. Notice the brown ring which forms. (See "Modern Chemistry," page 85, for another plan of testing the same.)

CHAPTER VII

THE HALOGENS

1. Hydrofluoric Acid, HF. — Warm a sheet of glass about 4 inches square by holding it some distance above a Bunsen burner, and drop upon it a few shavings of wax. When the melted wax is evenly distributed over the glass, allow it to cool. With a sharp pencil draw in the wax a design you wish to etch.

2. Put into a small lead dish, or evaporating dish, 2 or 3 g. of calcium fluoride, CaF_2, and cover with strong sulphuric acid. Put the prepared glass over the dish, wax down, and support upon a ring-stand. Warm very gently so as not to melt the wax. Allow it to remain 10 to 15 minutes. Remove the wax with a dull knife and a cloth. Note the odor of the gas, but be careful not to inhale much of it. Has it any color? Write the reaction, indicating the chemical changes.

3. Chlorine. Preparation. — Take every precaution to keep the gas out of the room; use the hood, if possible. Chlorine is usually collected by downward displacement (see page 13), but its escape into the room may be prevented better by collecting over a saturated solution of common salt. When you have filled as many bottles of chlorine as you need, let the delivery tube dip into a solution of sodium hydroxide. Why can you not collect chlorine over water? Why can you collect it over salt water?

4. Support a 100 cc. flask upon a ring-stand, and arrange to heat by water bath. (See instructions, page 15.) Put into the flask a few grams of manganese

dioxide, and add 25 to 30 cc. of strong hydrochloric acid. Attach a delivery tube and heat gently; collect six bottles of the gas.

5. Characteristics. Effect upon Burning Substances. — Into a bottle of chlorine put a burning match; results? Lower slowly a burning candle; results?

6. To compare Chlorine with Oxygen in Chemical Activity. — Into a second bottle of chlorine sift some finely powdered arsenic or antimony; results? Recall work in oxygen; did you find any substances that would burn without igniting? Put a small piece of phosphorus into a deflagrating spoon, and lower into a third bottle of chlorine; results? Compare with oxygen.

7. Chemism of Chlorine for Free Hydrogen. — Start a small hydrogen generator working, attach a bent jet, and when the air is all expelled (*see caution under hydrogen*) light the gas. Hold the burning jet in a bottle of chlorine; does it continue to burn? What change do you notice taking place? Why? When you see no further change occurring, pour a few cubic centimeters of water into the bottle and shake it. Test the solution with litmus paper; results? Why? Would a solution of chlorine do this? You can answer after trying the experiment given in section 9.

8. Chemism of Chlorine for Combined Hydrogen. — Saturate a narrow strip of blotting paper in hot turpentine, and insert by means of forceps in a bottle of chlorine; results? What has been set free? (Turpentine consists of carbon and hydrogen.) Explain results. If you do not perform this experiment, answer the questions from work in the class room.

9. Chlorine as a Bleaching Agent. — Suspend in a bottle of chlorine wet and dry pieces of blue litmus paper, and

give results in each case. Can you explain, reasoning from what you saw in section 8? Try a piece of moistened calico; result? Also a piece of newspaper, with some red ink written across the printing; explain results.

10. Put into one test-tube a few cubic centimeters of a solution of litmus; into a second, a solution of carmine or cochineal; into a third, a solution of copper sulphate. Into each pass a few bubbles of chlorine from the generator. What difference do you find in results? Explain, in accordance with facts learned in 8 and 9.

11. Make a note of the color, odor, density, combustibility, and chemical activity of chlorine.

12. The Chemical Action. — Put a small crystal of potassium chlorate, $KClO_3$, into a test-tube, add about a half cubic centimeter of hydrochloric acid, and warm gently. What gas is evolved? What gas have you prepared before from this solid by heating?

13. Into another test-tube put a very little powdered potassium dichromate, and heat with a little hydrochloric acid; can you detect the presence of any gas? What?

14. Into a third tube put 5 or 6 cc. of strong sulphuric acid, and add a few grams of manganese dioxide. Shake well, attach a delivery tube, and collect over water a test-tube full of the gas generated by heating. Test it for oxygen; results?

15. Put a half-gram of sodium chloride into a test-tube, add a little sulphuric acid, somewhat diluted; if necessary, warm gently; results? Blow across the top of the tube; results? The gas is hydrochloric acid. Any color?

16. Now combine 14 and 15; that is, put a little manganese dioxide and sodium chloride, mixed, into a tube, add some sulphuric acid, shake them together, and warm gently. Are the results the same as in 14? Or the same

as in 15? What gas is obtained? Do the following equations represent what you obtained in 14 and 15?

$$(14)\quad MnO_2 + H_2SO_4 = O + x.$$

$$(15)\quad NaCl + H_2SO_4 = HCl + y.$$

If so, explain what has probably happened in 16. Compare with other methods used in preparing chlorine, and see whether your conclusions are justified.

17. Hydrochloric Acid, HCl. Preparation. — How have you prepared the corresponding compound of fluorine? How have you prepared hydrochloric acid? Try in the same way a little potassium chloride and add sulphuric acid. Notice the odor. Blow across the tube; results? Do they agree with those previously obtained? What general plan might you suggest, judging from these two experiments, for preparing hydrochloric acid?

18. Preparation in Larger Quantities. — Put into a 200 cc. flask, *F*, a small handful of sodium chloride, and cover with strong hydrochloric acid. Fit the flask with a 2-hole rubber stopper and arrange the apparatus as shown in Fig. 23. Partly

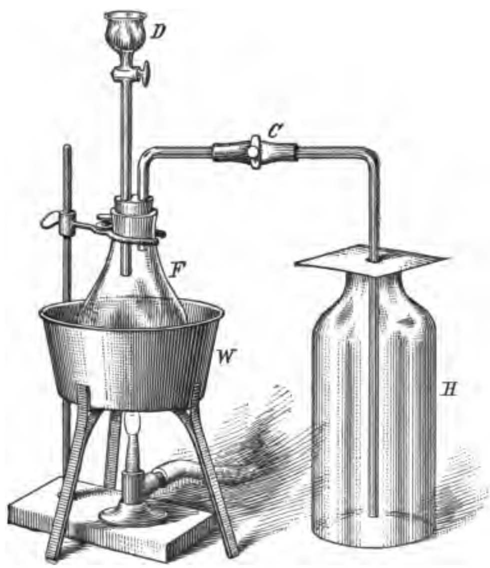

FIG. 23.

fill the funnel, *D*, with strong sulphuric acid, and support *F* upon a water bath. Collect the gas by downward displacement in *H*, which it is well to keep covered with a card, perforated to admit the delivery tube. Warm the bath and allow the sulphuric acid to drop very slowly into the flask. At first the evolution of gas will be irregular, but it will soon become constant, and may be controlled perfectly by regulating the flow of acid into the flask. Just before you are through collecting the gas, remove the flask from the hot water and close the stop-cock of *D*. A burette may be used instead of the funnel, if you have not the latter; or a small funnel with rubber connections and screw clamp. For some experiments the gas should be dried by passing through a bottle containing sulphuric acid.

19. Collect in this way four bottles of the gas. Note its color, its density, its effect upon moistened litmus paper, and how it is different from chlorine in this regard. Put a burning candle into it; results?

20. To a second bottle add a few cubic centimeters of water, and shake. Taste a drop of the solution; try litmus paper; results? Save contents.

21. Invert a third bottle, well filled, over water, and note results. Explain.

22. Moisten a stirring rod with ammonium hydroxide, and put into a bottle of hydrochloric acid gas; results? Why?

23. **Test for Hydrochloric Acid**. — Put a drop of the hydrochloric acid solution upon your table into a test-tube and add a little water. Now add a few drops of a solution of silver nitrate; results? To this add ammonium hydroxide, till alkaline; results? Still further, add nitric acid, till you can no longer smell the ammonia in the

tube; results? Repeat these tests, using the solution saved in 20 above; results?

24. Bromine. Preparation. — Use great care to keep the vapors out of the room. Make small quantities and use the hood. What have you found to be the result of treating manganese dioxide with sulphuric acid? Sodium chloride with sulphuric acid? (See page 40.) What, therefore, would you expect if you treated sodium or potassium bromide with sulphuric acid? Make an experiment to see whether your conclusions are correct; results? Is any hydrobromic acid formed? Blow across the top of the tube; results? Test also with ammonium hydroxide on a glass rod; results? Hold a moistened piece of red litmus paper inside the tube; result? Look up explanation of the difference between this gas and chlorine.

25. Put into a test-tube a few small crystals of potassium or sodium bromide, add a little manganese dioxide, and cover with strong sulphuric acid. Heat gently. What appears in the tube? Upon the sides? Is any hydrobromic acid obtained?

26. Characteristics of Bromine. — Attach a delivery tube and pass a few bubbles of gas into a cubic centimeter of a carmine solution. Results? Are you able to compare its bleaching power with that of chlorine? Note any physiological effects. Make a note, also, of any other characteristics learned by experiments in the class room.

27. Iodine. Preparation. — Determine whether you can prepare hydriodic acid by treating potassium iodide with sulphuric acid; test with ammonium hydroxide on a glass rod, also by blowing across the tube; results? It is a colorless gas. Is anything else obtained?

28. Put a few small crystals of potassium iodide mixed with manganese dioxide into a test-tube and add a little

strong sulphuric acid; warm gently; results? Is hydriodic acid formed now? Explain (see page 40).

29. Characteristics. — What is the color of iodine vapors? odor? effect on litmus? upon the fingers? Try to remove this by treating with ammonia; explain. Heat a small crystal of iodine in a test-tube; results? Did the crystal melt?

30. Ionized Halogens. — As we have seen chlorine, bromine, and iodine, they have been in the molecular condition; that is, a large number of molecules collectively, and have appeared under different colors. *Ionized*, as in solutions of sodium or potassium chloride, bromide, and iodide, in which, if dilute, the molecules of these compounds are largely broken up, or *dissociated*, into their constituent atoms, the same elements are colorless.

31. Put into a test-tube 1 cc. of a solution of potassium bromide and add to it a few bubbles of chlorine, or a few drops of chlorine water; results? Explain the action of the chlorine in accordance with the suggestions above. Try in the same way the effect of chlorine upon a solution of potassium iodide. Explain.

32. Test for Free (Molecular) Bromine or Iodine. — If not free, add to the solution, to be tested, a little chlorine, then about 1 cc. of carbon disulphide, and shake. Try thus solutions of potassium bromide and potassium iodide. Note results upon adding the chlorine, and after the next step; how alike? how different? Free iodine may also be tested with a starch solution. Try this, diluting the iodine solution till of a pale straw color; results?

33. Test for Ionic Chlorine, Bromine, and Iodine. — Make a solution of potassium bromide or iodide, and determine whether carbon disulphide or starch will show the presence of bromine or iodine ions; conclusion? Take

three test-tubes ; put into the first 1 cc. of a sodium chloride solution ; into the second, sodium bromide solution ; into the third, sodium iodide solution. To each add a few drops of silver nitrate solution. Compare results. To each now add a little ammonium hydroxide, and shake ; compare results. This test for ionic bromine and iodine should always be confirmed by converting into the free condition and testing as above.

CHAPTER VIII

BASES, ACIDS, AND SALTS

1. Bases. — Test a solution of sodium hydroxide with red litmus paper ; results ? Taste it. In the same way test a solution of potassium hydroxide, lime-water, and ammonia. What is the general result ?

2. Acids. — Add a drop of sulphuric acid to several cubic centimeters of water. What is the taste ? Effect upon litmus ? In same way try nitric, hydrochloric, and acetic acids ; results ? From work previously done state what is the result of treating a metal with an acid, as zinc or iron with sulphuric or hydrochloric acid.

3. Acidic and Basic Oxides. — What is an oxide ? Ignite a little sulphur in a deflagrating spoon and hold in a bottle for a few minutes. The gas is sulphur dioxide. Add a little water to the bottle, shake, and test with litmus paper ; what is the character of the solution ? In the same way try a *very* small piece of phosphorus ; results ? What kind of oxides are these, of sulphur and phosphorus ? They represent a large number of others. What other name is often given to such oxides ? Review work in oxygen, section 8.

4. Dissolve a small fragment of lime, CaO, in a little water in a test-tube. Try its effect upon litmus paper; results? Try also barium oxide, BaO; results? What is the nature of these oxides? What compounds does this class of oxides form with water?

5. **Neutralization. Salts.** — To 1 cc. of a moderately weak solution of sodium hydroxide add one drop of a solution of phenol phthalein; results? Now add drop by drop hydrochloric acid until some decided change takes place; what? Again add sodium hydroxide slowly; results?

6. Put into an evaporating dish 20 cc. of dilute hydrochloric acid and add one drop of phenol phthalein solution. Next, with constant stirring, add a solution of sodium hydroxide until the faintest pink color begins to show. What does this indicate? Then add one drop of hydrochloric acid. Evaporate the solution nearly to dryness, or until a drop, cooled on the end of a stirring rod, crystallizes. Let the solution cool, filter out the crystals, and wash with a *very little* cold water. Dry them. What is the appearance? Taste? Effect of a solution of the crystals upon litmus? Two respects in which this *salt* differs from either the base or the acid used in preparing it? Could you dry moist sodium hydroxide in the air? What other salts have you prepared in a similar way? (See Ammonia.)

7. **Normal Salts.** — What is a normal salt? Put 25 cc. of dilute sulphuric acid into an evaporating dish and add one drop of phenol phthalein. Allow a solution of sodium hydroxide, after reading the burette, to run slowly into the acid, stirring constantly, until *exactly* neutralized. Note the amount of solution used. Evaporate till the solution will crystallize upon cooling. (While waiting,

proceed with the next work.) Filter out the crystals, dry, and preserve for future examination.

8. Acid Salts. — What is an acid salt? Put 25 cc. of sulphuric acid of the same strength as that used in the preceding experiment into an evaporating dish, and add to it exactly half as much sodium hydroxide solution as you used before. Evaporate to point of crystallization, filter out, and dry. Compare the two salts, thus prepared, in appearance, taste, and shape of crystals. Will a solution of either one affect litmus paper? Determine whether both contain water of crystallization.

9. Ionization. — What is an ion? What is meant by ionization? Refer to comparison of the halogens; also to the collection of chlorine over salt water. Why is chlorine not very soluble in a saturated solution of sodium chloride, but soluble in pure water?

10. Dilute solutions of ionizable substances are much more completely dissociated than strong solutions: thus, concentrated hydrochloric acid (35 per cent) is only ionized to about 14 per cent, while a solution about one-tenth as strong is ionized over 75 per cent. Therefore in a dilute solution of hydrochloric acid we have positive ions, *kathions*, of hydrogen; negative, or *anions*, of chlorine, and undissociated hydrochloric acid; also, in the water, kathions of hydrogen, anions of *hydroxyl*, HO, and undissociated molecules of water. Sodium hydroxide dissociates to about the same extent as hydrochloric acid.

11. If we put together these two substances, both in dilute solutions, what ions would be present? What four undissociated substances? If the acid and base above named exactly neutralize each other, what happens as you boil down the aqueous solution? What is the effect of completely expelling the water?

CHAPTER IX

CARBON AND A FEW COMPOUNDS

1. Charcoal. — How is it prepared? Heat to redness a small piece of charcoal and hold it under water for a moment. Does it float? Try a piece which has not been thus heated; results? Why? If you cannot explain, hold the piece, which was not heated, under water and observe what happens. What was the effect of the heat?

2. Proof that Charcoal is an Absorbing Agent. — Put into a 100 cc. flask 10 or 15 cc. of water into which hydrogen sulphide has been passed; add about a teaspoonful of powdered charcoal or bone-black, cork the flask, and shake vigorously. Is the odor still noticeable? Explain. Try in the same way water rendered offensive by decomposing vegetable or animal matter; results?

3. Practical Value of this Characteristic. — Fit to a funnel a filter paper and fill about halfway with powdered bone-black. Pass through it slowly a solution of brown sugar, molasses, or water colored with litmus or other organic pigments; results? What practical uses are thus illustrated?

4. Carbon, a Good Reducing Agent. Proof. — What is a reducing agent? (See "Modern Chemistry," pages 35 and 139.) With the handle of your steel forceps, bore a hole in a stick of carbon near one end, and fill it nearly full of lead oxide, PbO. Heat it strongly with the *reducing* flame. (See page 12 for directions.) Results? Explain.

5. Destructive Distillation. — Put into a test-tube a few grams of powdered soft coal, attach a cork and delivery tube. Support upon a ring-stand and heat, at first gently and then strongly, for several minutes. What is driven off from the coal? Test its combustibility; result? Note the odor. Pass it into a beaker of water for some time; test the water with litmus; result? When the gas ceases to come off, break open the tube and examine the contents; what is the residue? Why is this process called destructive distillation?

6. Fractional Distillation. — (For special students.) Arrange apparatus as shown in Fig. 24. A retort may be used if no distilling flask is at hand. *A* is a distill-

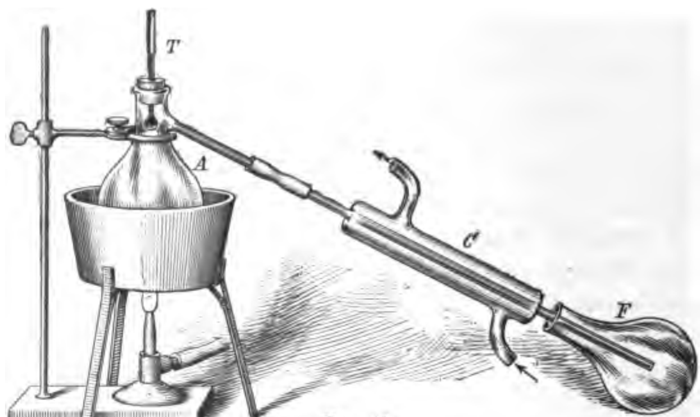

FIG. 24.

ing flask, supported by a clamp from the ring-stand, upon a water bath. *T* is a thermometer, passing through the cork to the level of the side neck. *C* is a Liebig condenser, *F* the receiving flask. Make connections with the water supply so that the current shall be in the direction of the arrows.

7. Put into the flask, *A*, 50 cc. of water and 100 cc. of common alcohol, gradually heat the water bath and note the temperature at which the mixture just begins to boil. Collect what comes over while the thermometer is rising 10 degrees. Collect a second fraction during the next rise of 10 degrees; collect also a few cubic centimeters as a third fraction. Examine the three fractions, and note wherein they differ. The experiment is more striking if about 25 cc. of fusel oil is used instead of the water above, and then fractions of about 35 cc. each be collected. Test a portion of each by adding a little water. Fusel oil is not soluble in water, and you will thus obtain evidence as to the amount in each fraction; results?

Carbon Dioxide, CO_2

8. **Preparation.**—Arrange apparatus as for generating hydrogen, but collect the gas by downward displacement. Put into the flask several small lumps of marble, cover with dilute hydrochloric acid, and collect five bottles of the gas.

9. **Characteristics.**—Has the gas odor? Color? Test its combustibility and power of supporting combustion in the usual way; results? In a second bottle try ignited magnesium ribbon; results? What two products do you see? Explain and compare with nitric oxide. Invert a third bottle over water; is there any evidence that the gas is soluble? If not satisfactory, add a little water to the bottle, and shake it. Test the water with litmus paper; conclusion? Is carbon dioxide an acidic or basic oxide? (Refer to work in oxygen and page 45.)

10. Pour some clear lime-water into a small bottle or test-tube of carbon dioxide, and shake; results? This is

the usual test for this gas. Try pouring the gas from one bottle into another; determine whether you have succeeded; conclusion?

11. Natural Sources of Carbon Dioxide. — Put into a flask a little clear lime-water, and by means of a tube blow the breath from the lungs through it; results? Lower a small candle into a bottle of carbon dioxide, and allow it to burn a few minutes, or until it goes out. Pour a little clear lime-water into the bottle, and shake; results? Can you name a third natural source of carbon dioxide?

12. To find the Per Cent of Carbon Dioxide in a Compound. — (Adapted from Fresenius.) Arrange two very small flasks or bottles, of 50 cc. capacity, as shown in Fig. 25. At *E* is a screw-clamp for closing the rubber tubing air tight. Powder the carbonate to be analyzed, weigh out a small quantity of it, about 1 g., carefully, and put into the bottle *M*. Add a little water. Fill *N* about half full of strong sulphuric acid. Make *E* air tight, and weigh the whole apparatus. The corks should be of rubber. By suction partially exhaust the air in *N*, then allow the air to reënter. The acid will be partly forced over into *M* and will decompose the carbonate. It may be necessary to force the acid over a second time. When the action has entirely ceased, open the clamp *E*, and by suction remove the carbon dioxide from the two bottles. You can tell by the taste when you have done this. Cool the apparatus, and weigh again. The loss in weight represents

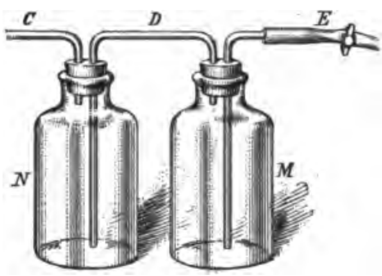

FIG. 25.

the carbon dioxide expelled. Determine what percentage it formed of the original compound.

13. To find Weight of One Liter of Carbon Dioxide. — (For special students.) Arrange apparatus as shown in Fig. 26. *A* is a flask for generating carbon dioxide; *B* is a wash-bottle containing sulphuric acid. *C* should have a capacity of 200 to 300 cc. Find the weight of this flask together with the perforated cardboard, *M*, used as

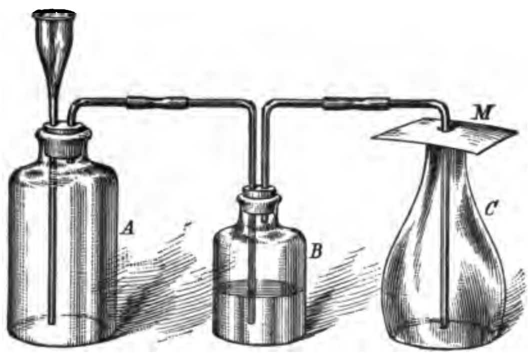

FIG. 26.

a cover. Call this weight *a*. Put into the generating flask about 25 g. of finely powdered marble, and add enough water to seal the thistle tube. Next add dilute sulphuric acid to the marble. (Unless the marble is finely pulverized, sulphuric acid will not be satisfactory. Its use has some advantages, hence it is suggested here.) After several minutes weigh the flask and cardboard; put the flask into place and again pass in the gas and weigh. Do this as long as it gains in weight. Call this weight *b*.

14. Determine carefully the capacity of the flask by filling with water. Determine what this volume of air would weigh at the temperature and pressure of the room,

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weight c ("Modern Chemistry," page 96). One liter of air under standard conditions weighs 1.293 g. Then

$$a - c = \text{empty flask} + \text{cardboard},$$
$$b = \text{flask} + CO_2 + \text{cardboard},$$
$$b - (a - c) = \text{wt. of } CO_2.$$

From this determine the weight of 1 l. of carbon dioxide; what is your per cent of error?

HYDROCARBONS

15. **Marsh Gas, CH_4.** — Describe from work in class room.

16. **Ethylene, C_2H_4.** — This may be prepared by treating alcohol with sulphuric acid, and heating. The following reaction is usually written to represent what takes place: —

$$C_2H_5OH + H_2SO_4 = C_2H_4 + H_2SO_4,H_2O$$

Really, an ester of sulphuric acid forms as an intermediate product, which is unstable and breaks down into ethylene and sulphuric acid, thus: —

$$C_2H_5OH + H_2SO_4 = \left.\begin{matrix} C_2H_5 \\ H \end{matrix}\right\} SO_4 + HOH =$$
$$C_2H_4 + H_2SO_4 + H_2O.$$

If the gas is not prepared in the laboratory, describe from work in class room.

17. **Acetylene, C_2H_2.** — Put a small lump of calcium carbide, CaC_2, into a test-tube, and add a little water. Note the vigor of the action, color of the gas, odor, combustibility, substance remaining in the tube. Write the reaction. Fill a test-tube about a quarter full of water, and invert it over a dish of water. Pass in acetylene

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until the tube is full. Lift it from the water, and immediately touch a lighted match or candle; results? Conclusions as to a mixture of air and acetylene?

18. A fairly good acetylene generator may be made by arranging the apparatus as for making hydrogen. Substitute a dropping funnel for the thistle tube. Put the carbide into the flask, and let the water drop slowly upon it.

CAUTION. — Never bring a flame near the opening of a delivery tube from an acetylene generator until a sample of the gas collected in a test-tube burns quietly when ignited.

19. A better generator for the class room is described in "Modern Chemistry." Let the student write up any further experiments made by the instructor.

CHAPTER X

SOME IMPORTANT LAWS

1. The Law of Definite Proportions. — We have had this illustrated already, but a few additional experiments will emphasize it. Support two burettes upon a stand; fill one with dilute hydrochloric acid, the other with sodium hydroxide. Carefully take the reading of each, after allowing the solution to run down and fill completely the tube below the stop-cock.

2. Counterpoise with shot or weigh accurately a small evaporating dish, and run into it 10 cc. of the sodium hydroxide solution. Add one drop of phenol phthalein or methyl orange solution. If more convenient, litmus may be used. Now run in slowly the acid, stirring constantly, until the alkali is exactly neutralized. Upon a

water or sand bath evaporate the solution to dryness. Do not let the liquid boil. When perfectly dry, cool and weigh. Determine weight of sodium chloride obtained.

3. While waiting for the above to evaporate to dryness, repeat the experiment, using the same amount of alkali and twice as much acid. Omit the use of the *indicator*. Repeat, using 5 cc. of alkali and sufficient acid to neutralize. Tabulate all results, thus: —

wt. of dish + NaCl ——	NaOH used ——,	
wt. of dish ——	HCl used ——,	
wt. of NaCl ——.		

What ratio between the first and second experiments? Between the second and third? Explain.

4. Other Proof. — Part of the students may make the following experiments while others are at work upon the preceding. Counterpoise, or weigh accurately, an evaporating dish; put into it about a half gram of sodium carbonate. Weigh carefully. Cover with water, and add dilute hydrochloric acid slowly. Keep the dish covered with a glass so as to avoid loss through the effervescence. Thus add the acid till the sodium carbonate is all decomposed; wash off the cover-glass into the dish, and evaporate slowly to dryness upon the water or sand bath. When perfectly dry, cool and weigh. Tabulate results. While waiting for this solution to evaporate, go on with the following. Repeat, using twice as much sodium carbonate as before. Tabulate results. Repeat, using three times as much sodium carbonate as in the first experiment. Tabulate results. What ratio exists between the first and second experiments? Between the first and third? Explain.

5. Combining Weight of Copper. — Put into a beaker $2\frac{1}{2}$ g. of bright copper turnings, and dissolve slowly in

nitric acid. Take care to prevent loss by spurting. Rinse off the cover-glass, and transfer the solution to an evaporating dish. Rinse out the beaker into the dish, and evaporate the solution to dryness. When perfectly dry, remove the copper nitrate from the water bath and heat slowly to dull redness, until you have a black powder. Cool and weigh. This black powder is copper oxide,—

$$Cu(NO_3)_2 = CuO + 2\,NO_2 + O.$$

(See work with nitrates.) Determine the amount of oxygen which has combined with the copper by subtraction. Assuming 16 to be the combining weight of oxygen, find that of copper, thus :—

comb. wt. of O : comb. wt. of Cu : :
wt. of O found : wt. of Cu used;

or, $16 : x :: M$ (found above) : $2\frac{1}{2}$ g.

Compare value of x with weight of copper given in any table of the elements.

6. To find the Strength of an Acid Solution.—Suppose we are required to determine how much acetic acid there is in a sample of vinegar. Find the weight of 10 cc. of the vinegar. If dark-colored, dilute with water till it is moderately clear. Add a drop or two of phenol phthalein, and then run in from a burette a decinormal solution of potassium or sodium hydroxide until the pink is just fixed. Suppose you use 4 cc. of the alkali solution; determine the amount of acetic acid from the following equation:—

$$\begin{array}{ccccccc} KOH & + & HC_2H_3O_2 & = & KC_2H_3O_2 & + & H_2O, \\ 56 & + & 60 & = & 98 & + & 18. \end{array}$$

That is, 56 g. of caustic potash would neutralize 60 g. of acid; or, as in a decinormal solution there are 5.6 g. in a

liter, 1000 cc. of caustic potash solution would neutralize 6 g. of acetic acid. Then —

$$1000 : 6 :: 4 : x,$$

in which x is the amount of acid in 10 cc. of the vinegar. From this find the amount in 1000 cc. or determine at once the percentage of acid, knowing the weight of the 10 cc. of vinegar taken.

7. **To find the Amount of Alkali in a Solution.** — Prepare a solution of oxalic acid by dissolving 6.3 g. of the crystals in 1 l. of pure water. When oxalic acid reacts with sodium hydroxide, it may be represented by the following: —

$$H_2C_2O_4, 2\ H_2O + 2\ NaOH = Na_2C_2O_4 + 4\ H_2O,$$

$$\text{or, by weight, } 126 \quad + \quad 80 \quad = \quad 134 \quad + \quad 72,$$

or 6.3 g. oxalic acid would neutralize 4 g. sodium hydroxide.

8. Measure out accurately into a beaker 25 cc. of the alkali solution to be tested, and add one drop of phenol phthalein or some other indicator. From a burette run in the solution of oxalic acid slowly till neutralization is effected. Note the number of cubic centimeters used. From this we can determine the amount of alkali. By the equation above, we see that 6.3 g. of acid combine with 4 g. of the alkali. But 6.3 g. of acid are contained in 1000 cc. of the solution. Then —

$$1000 \text{ cc. acid} : 4 \text{ g. NaOH} :: m \text{ (no cc. acid used)} : x.$$

But x is the amount of sodium hydroxide in 25 cc.; from this find the amount in 1 l., 1000 cc. For other problems of similar character, see "Modern Chemistry," page 169, or any work in volumetric analysis.

CHAPTER XI

SULPHUR AND SOME COMPOUNDS

1. Various Forms of Sulphur Crystals. — Add a little powdered brimstone to 1 or 2 cc. of carbon disulphide in a test-tube, and shake. When the sulphur has dissolved, pour out the solution upon a watch glass, and set it to crystallize in as cool a place as possible. Later, examine the crystals with a lens, and note their form. Make a drawing of one of the crystals.

2. If this experiment is not done in the laboratory, write up from work in the class room. Fill a small sand crucible nearly full of sulphur, and heat till the contents are melted. Let them cool slowly, and when a crust has formed over the surface, break it and pour out the molten contents. In a few minutes break the crucible and examine the crystals. Compare them with those obtained in section 1.

3. Plastic (Amorphous) Sulphur. — Put a few grams of sulphur into a test-tube, and gradually heat to near the boiling-point. What various changes take place? Pour out the thin liquid into a dish of cold water. Examine the sulphur. How is it different from any forms previously examined? Test the solubility of a small portion of it in carbon disulphide. Put the remainder away for a few days; again examine it, and note changes.

4. Name two other forms of sulphur besides those you have prepared. How are they obtained?

Hydrogen Sulphide

5. **Method of Preparing.** — Use the hood and make very small quantities. Put a small piece of ferrous sulphide, FeS, into a test-tube, and cover with water. Add a few drops of strong sulphuric acid; no heat is necessary. Action begins at once. Write the reaction.

6. **Characteristics.** — Note the color and odor of the gas. Insert a cork and jet. After a moment or two of vigorous action, light the jet; how does it burn? What is the odor while burning? Hold a cold beaker over the jet; what condenses on the beaker? What are the two products of burning hydrogen sulphide — one seen on the beaker, the other detected by the odor? Write the reaction.

7. Pass hydrogen sulphide into water a few minutes, and test the solution with litmus paper; results?

8. **Reducing Action of Hydrogen Sulphide.** — (See page 49, and "Modern Chemistry," pages 35 and 139.) Put into one test-tube 2 or 3 cc. of potassium chromate solution, and add a few drops of hydrochloric acid; into another test-tube put 2 or 3 cc. of dilute ferric chloride solution: through each pass hydrogen sulphide for several minutes. What change takes place in the appearance of each? The chromate, K_2CrO_4, has lost its oxygen and become a chloride, while the ferric solution, $FeCl_3$, has become a ferrous solution, $FeCl_2$. Test the ferrous ions by adding a few drops of ammonium hydroxide; results? To 1 cc. of a ferrous sulphate solution add ammonia in the same way; results? Have you really obtained a ferrous solution by using the hydrogen sulphide?

9. **Use of Hydrogen Sulphide.** — Its chief use is in analytical work, which may be shown in an experiment.

Put into a flask a few cubic centimeters of lead nitrate solution and about an equal amount of dilute ferric chloride solution. Disregard any slight precipitate that may form. Pass hydrogen sulphide through the solutions until the odor continues perceptible in the flask after shaking. Filter out the precipitate. Let the gas again bubble through the clear filtrate for a moment; if no precipitate forms, it no longer contains any lead ions. Test it for iron in the ferrous condition, as you did in the preceding section. Have you succeeded in separating the lead and iron salts? These two salts represent a large class which may be separated in a similar manner.

Sulphur Dioxide

10. Preparation. — Put into a small flask a few grams of copper turnings, and cover with strong sulphuric acid. Attach a cork with delivery tube, and arrange to collect the gas by downward displacement, as you did carbon dioxide. Support the flask upon a ring-stand with wire-gauze protection, and heat until vigorous action begins. Fill three bottles with the gas, keeping them covered till ready for use. Remove the heat, but leave the flask in position; allow the delivery tube to dip just under the surface of some water in a beaker, so as to keep the gas out of the air.

11. Characteristics. — Put a little cochineal or litmus solution into a test-tube, and allow a current of sulphur dioxide to pass into the solution; results? Compare with chlorine, if you can. Note the color and odor of the gas. What can you say of its comparative density? Try its effect upon litmus paper; is it an acidic or a basic oxide? Test the water into which the gas has been flow-

ing; is it acidic or basic in its reaction? Test the combustibility of the gas; does it support combustion? Invert a second bottle over a pan of water; results? Explain. In a third bottle suspend a moistened piece of calico and some colored paper; allow them to remain some time; results? How does the action compare with that of chlorine?

12. Sulphur Dioxide as a Reducing Agent. — (Refer to "Modern Chemistry," pages 35, 49, and 179 in this book.) Put into a test-tube a few copper turnings, and fit to it a cork and delivery tube. To the copper turnings add a little nitric acid, and pass the gas evolved into a bottle by downward displacement. What gas is first produced? What have you in the bottle? Why do you think so? Cover the bottle with a perforated cardboard and pass into the brown fumes a current of sulphur dioxide; results? Did the brown fumes escape into the air? Remove the delivery tube and take the cover from the bottle; hold a white paper behind it; do you notice anything that helps you explain the action of the sulphur dioxide? What has it really done?

13. Put into one test-tube a little potassium permanganate solution, and into another a little potassium chromate or dichromate solution. Pass into each a current of sulphur dioxide until some decided change occurs; results? Compare with hydrogen sulphide (page 60). Determine whether you can reduce ferric salts to ferrous salts by sulphur dioxide. (See hydrogen sulphide.)

14. The Chemical Action. — Compare work with hydrogen in kind and strength of acid used. In temperature at which action occurs. In gaseous products obtained. What two steps probably occur in preparing sulphur dioxide? (Refer to "Modern Chemistry," page 178, and to

nitric oxide, page 36, of this book.) What characteristic of hydrogen is illustrated in these two instances?

15. Further Study of the Chemical Action. — (For special students.) Pour the contents of the flask used in generating sulphur dioxide into 100 cc. of water; after a little time filter, if necessary, to remove any sediment. Evaporate slowly upon a water or sand bath until a drop of the solution upon a glass rod will deposit crystals. Allow it to cool, drain off any liquid from the crystals, put them upon a filter paper in a funnel, and wash with a very little cold water. Dry and preserve them in a specimen tube. Describe them.

16. Dissolve one of the crystals in a little water, and test for copper ions by adding ammonium hydroxide slowly. The usual result is a pale blue precipitate at first, which soon dissolves in excess of ammonia, forming a deep blue solution. What are your results? Dissolve another crystal in water, and test for the negative ion. To do this, add a few drops of hydrochloric acid, then a little barium chloride solution. A white precipitate indicates the sulphate ion, SO_4; results? What is the blue compound? Write the reaction as far as you have worked it out.

$$Cu + x\,H_2SO_4 = \text{——} + \text{——} + y.$$

FIG. 27.

17. Liquefaction of Sulphur Dioxide. — (For special students.) Arrange to generate the gas in a flask of about 200 cc., and dry it by passing through a bottle partly filled with sulphuric acid. Next conduct it into a spiral glass tube sur-

rounded by a mixture of ice and salt in a quart can, and let it pass out the bottom through a rubber cork which fits tightly. With short rubber connections join the spiral tube to a tube reaching to the bottom of a test-tube, also surrounded by a freezing mixture. In this way several cubic centimeters may be liquefied in a few minutes. If the student desires, he may preserve a specimen in a strong glass tube. For method of sealing, refer to Chapter I, section 6.

SULPHURIC ACID, H_2SO_4

18. Preparation. — Support two 100 cc. flasks upon ring-stands, as shown in Fig. 28. Protect *A* by wire gauze. Arrange delivery tubes passing to the bottom of *C*, a

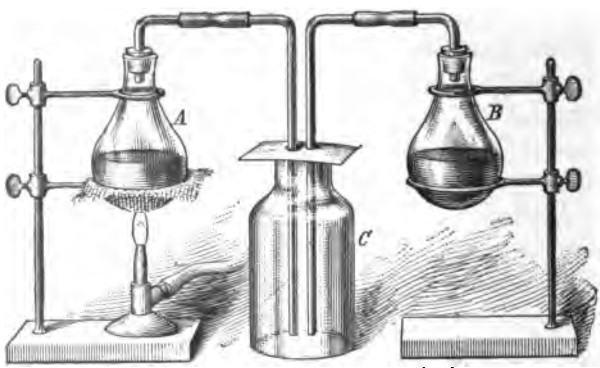

FIG. 28.

200 cc. bottle or flask, and let them pass through a doubly perforated cardboard. Arrange to generate sulphur dioxide in *A*, and nitric oxide in *B*. First fill *C* with nitric oxide; explain the presence of the brown fumes. Then pass in sulphur dioxide; explain the change

which occurs. Finally, add a little water, and shake. Test the solution with barium chloride solution and hydrochloric acid, as in section 16; results? Conclusion?

19. The Chemical Action. — Write a reaction expressing what happened when the nitric oxide entered the bottle. Another, showing what happened when the sulphur dioxide came in contact with the nitrogen tetroxide. Write a third, indicating the action of the water. How does the commercial method of preparing sulphuric acid differ from this? In answering, state the source of the sulphur dioxide and of the nitric oxide; also, how the water is added. (See "Modern Chemistry," page 181.) What is the anhydride of sulphuric acid?

20. Sulphurous Acid, H_2SO_3. — How have you prepared this? (Refer to section 10.) What is the anhydride of sulphurous acid?

21. To determine the Molecular Weight of a Gas. — (For special students.) Suppose it is desirable to determine experimentally the molecular weight of a gas, as, for example, of sulphur dioxide. Arrange apparatus and determine the weight of 1 l. of sulphur dioxide, as you did for carbon dioxide. (See figure, page 53.) Prepare the sulphur dioxide in the usual manner. Having determined the weight of a liter, you can find the molecular weight from the following formula: —

$$\frac{\text{mol. wt.}}{2} \times .0896 \text{ (wt. 1 l. H)} = \text{wt. 1 l. of the gas.}$$

Why is this equation true? (See "Modern Chemistry," pages 200 and 201.)

CHAPTER XII

PHOSPHORUS AND SOME COMPOUNDS

1. CAUTION. — Probably all experiments with phosphorus should be made by the instructor in the class room. If the classes are very small, so as to receive close supervision, the following experiments may be attempted in the laboratory. Let the greatest care be observed, as phosphorus burns are very serious.

2. Forms of Phosphorus. — Examine and compare the *ordinary* and the *red*, or *amorphous* varieties. What differences do you notice? Expose a minute quantity of each to the air. Wherein do they behave differently? To what two varieties of sulphur do they correspond? Put a piece of ordinary phosphorus, half as large as a pea, into about 1 cc. of carbon disulphide, and shake; results? How like sulphur in this particular? Pour upon a filter paper lying flat on a ring-stand; watch the paper a minute; results? Explain. From work in class room, note here the reaction of the halogens with phosphorus.

3. Hydrogen Phosphide, Phosphine, PH_3. — The preparation of this gas is not suitable for a laboratory experiment. From work in the class room, draw apparatus used, and describe method and results in detail. Describe the gas.

4. Oxides and Acids of Phosphorus. — Where have you seen the oxides of phosphorus prepared? Ignite a very small piece of phosphorus, and quickly invert over it a dry beaker or bottle of about 300 cc. capacity. Examine

the deposit upon the inside of the beaker. Quickly add to it a few cubic centimeters of hot water. Do you hear any sound as you pour in the water? Test the solution with litmus; results? Assuming that the white compound was mostly the pentoxide, P_2O_5, is it an acidic or basic oxide? What was formed when the water was added?

CHAPTER XIII

ACIDIC RADICALS, OR ANIONS

1. A number of these anions have been met with already, and their reaction with certain cathions has been noticed. For the sake of convenience and review, they are here summarized. If the acidic radical to be determined is in a solid, certain preliminary steps may aid very materially; but it must be remembered that these are merely preliminary and must be followed by other tests.

2. Preliminary Work. — Let the instructor furnish the student with a known sample of each of the following, in the solid form, and let the student verify the statement regarding it. Put a small portion of the sample into a test-tube, and add a little strong sulphuric acid. If necessary, warm gently. Notice whether effervescence occurs, and the color and odor of the gas, if any is evolved.

a. Acetates. — Give an odor like vinegar, no color.

b. Bromides. — Brown gas, offensive odor, irritating, especially to the eyes.

c. Carbonates. — Rapid effervescence, colorless gas, no odor; verified by passing gas into clear lime-water. (See carbon dioxide.)

d. Chlorides. — Colorless gas, very irritating; white fumes with ammonia, or if moist air from the lungs is blown across the top of the tube.

e. Iodides. — Violet gas.

f. Nitrates. — Colorless gas, irritating odor.

g. Nitrites. — Brown gas, irritating, but not so offensive as that of bromides.

h. Sulphates. — No results.

i. Sulphides. — Colorless gas, odor like decomposing eggs.

j. Sulphites. — Colorless gas, suffocating, odor of burning sulphur; turns drop of potassium dichromate solution on strip of paper green.

k. Thiosulphates. — Same as sulphites.

3. **Verification.** — Having learned by preliminary steps something as to the nature of the anion, you must now verify your previous conclusions by special tests. It is often necessary to know the cathion before the anion may be readily determined; but for the present work the instructor will furnish such exercises as will give little trouble, though studied in a reverse order.

4. After testing a small portion of the unknown substance by the preliminary tests, dissolve another part in a little water and proceed by the special test that you think the indications call for.

a. Acetates. — To the solution add a little ferric chloride solution and boil. Results: a deep reddish brown color, which disappears upon adding hydrochloric acid.

b. Bromides. — Add a few drops of chlorine water, and proceed as in section 32, page 44.

c. Carbonates. — Lime-water test.

d. Chlorides. — Add a little silver nitrate solution. A curdy, white precipitate forms, very soluble in ammonium hydroxide, which reprecipitates upon adding nitric acid.

e. Iodides. — (See work in iodine, comparison of the halogens, page 44.)

f. Nitrates. — Drop a crystal of ferrous sulphate into the unknown solution and pour carefully down the side of the test-tube a little strong sulphuric acid. A brown ring forms about the crystal.

g. Nitrites. — Add a solution of ferrous sulphate; the whole solution turns brown to black, according to the strength.

h. Sulphates. — Acidulate with hydrochloric acid, then add barium chloride solution. A heavy white precipitate, insoluble in any acid, forms.

i. Sulphides. — Odor is usually sufficient indication. Hold over the tube a paper moistened with a solution of lead acetate. This will turn black.

j. Sulphites. — Add strong hydrochloric acid to the solution. Results: emission of sulphur dioxide *without* any precipitate.

k. Thiosulphates. — Test as for sulphites. Results: evolution of sulphur dioxide, *with* a white precipitate of sulphur, becoming somewhat yellow if the solution is moderately strong.

5. **Exercise.** — Let the instructor furnish a considerable number of unknown salts to be tested as above for the acid radicals (anions).

CHAPTER XIV

THE ALKALI METALS

1. **Sodium. Characteristics.** — From work done in the class room, write a description of the metal: color, density compared with water, hardness. How is it affected by the air? Behavior when dropped on water, cold and hot? What is the reaction of the residual water upon litmus? What property has it acquired? Put a small piece of sodium upon wet blotting-paper, and when the molten metal has become bright like silver, let it roll off upon the floor; results? Explain.

Compounds of Sodium

2. **Caustic Soda, Sodium Hydroxide, NaOH. Preparation.** — Put into a beaker about 10 g. of sodium carbonate, dissolve in 50 cc. of water, and heat on the sand bath to boiling. Into another beaker put about 3 g. of lime,

[illegible]

[illegible]

[illegible]

means of an
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t the gas will
rapid. Keep
nty-four hours,
rt of the prep-
ting in with the
nning, as much

e, and drain them
a plug of mineral
r. Do this till the
es. (See page 68.)
his method of prepa-

water, the normal or
the acid (bicarbonate)
rbonate? If you have
vert half of it into the
the two samples in well-
all the reactions that take
aring the acid and normal

Purification. — Dissolve 25 g.
85 cc. of hot water, and imme-
oxide, made by putting about
ter. Next, add barium chloride
the clear salt solution will no
ate with it. After filtering, add
g. of sodium carbonate, and then
acid till no longer alkaline. Boil
stallization, and cool. Filter out the
with a very little water, and dry with

CaO, and add enough water to reduce to the consistency of cream. Add this milk of lime to the boiling solution of sodium carbonate, and continue to boil several minutes. Allow the precipitate in the solution to settle, and decant the *clear* liquid.

3. **Estimation of the Alkali Prepared.** — Into an evaporating dish, the weight of which is known, measure out half of the clear solution decanted, and boil slowly to dryness. While waiting, proceed with the next work.

4. **Titration of the Alkali.** — Take half of the remaining portion of the decanted liquid, add to it sufficient pure water to make about 30 cc., and titrate half of this with oxalic acid solution. (Refer to work in Chapter X, sections 7 and 8.) Make a second titration, and calculate the total amount of alkali in the half thus used. When the other half has boiled dry, cool and weigh quickly. How do the results by weighing and by titration agree?

5. **Sodium Bicarbonate (Acid-carbonate), $NaHCO_3$.** — (For special students.) Put 50 cc. of strong ammonium hydroxide into a small flask or bottle, and add 15 g. of common salt, finely powdered. Cork the bottle and shake for several minutes, or until the salt is all dissolved. Keep the solution as nearly ice cold as possible.

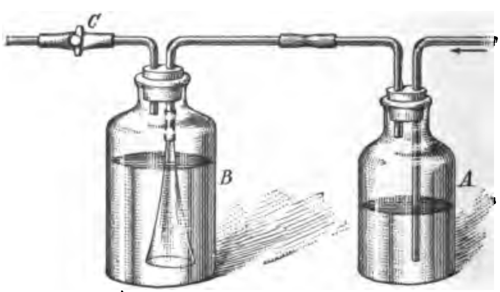

Fig. 29.

Next, arrange apparatus as shown in Fig. 29; connect with a Kipp generator. Let the carbon dioxide pass through a little water in the wash-bottle, *A*; intro-

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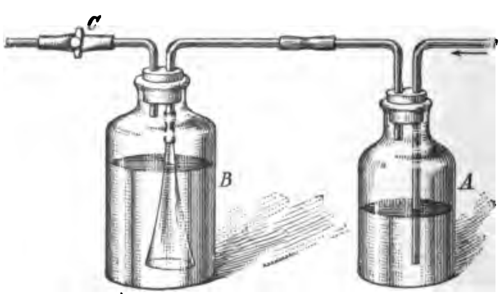

FIG. 29.

CaO, and add enough water to reduce to the consistency of cream. Add this milk of lime to the boiling solution of sodium carbonate, and continue to boil several minutes. Allow the precipitate in the solution to settle, and decant the *clear* liquid.

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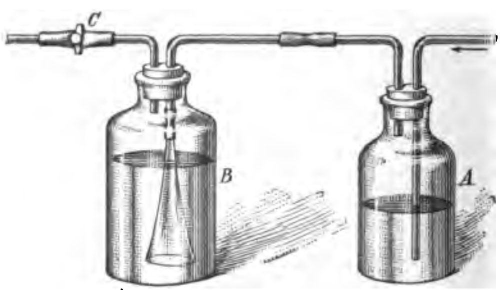

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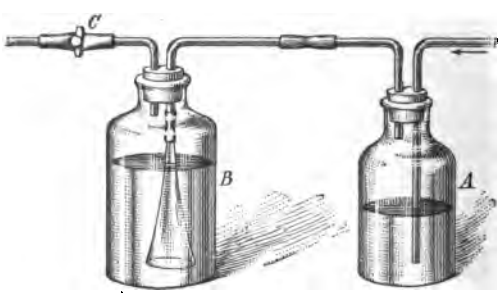

FIG. 29.

Next, arrange apparatus as shown in Fig. 29; connect with a Kipp generator. Let the carbon dioxide pass through a little water in the wash-bottle, *A*; intro-

duce the gas into the solution in *B* by means of an adapter, that is, a funnel-shaped tube. Why? Except at intervals, keep the clip at *C* closed so that the gas will be under pressure and the absorption more rapid. Keep *B* as cool as possible. Pass the gas for twenty-four hours, or as long as any crystals form. This part of the preparation may be hastened somewhat by putting in with the salt and ammonia solution, at the beginning, as much ammonium carbonate as will dissolve.

6. Remove the crystals from the bottle, and drain them in a funnel upon a platinum cone, or a plug of mineral wool. Wash with a little ice-cold water. Do this till the wash water gives no test for chlorides. (See page 68.) You now have sodium bicarbonate; this method of preparation is called the *Solvay* process.

7. Which is the more soluble in water, the normal or the acid carbonate? How may the acid (bicarbonate) be converted into the normal carbonate? If you have enough of the bicarbonate, convert half of it into the normal carbonate, and preserve the two samples in well-sealed specimen tubes. Write all the reactions that take place in this method of preparing the acid and normal sodium carbonate.

8. **Sodium Chloride, NaCl. Purification.**—Dissolve 25 g. of crude rock salt in about 85 cc. of hot water, and immediately add calcium hydroxide, made by putting about 0.5 g. lime into a little water. Next, add barium chloride solution until a little of the clear salt solution will no longer form a precipitate with it. After filtering, add a solution of about 1 g. of sodium carbonate, and then dilute hydrochloric acid till no longer alkaline. Boil down to point of crystallization, and cool. Filter out the crystals, wash them with a very little water, and dry with

duce the gas [illegible]
adapter, that is a [illegible]
at intervals, keep the [illegible]
be under pressure and the [illegible]
B as cool as possible. [illegible]
or as long as any [illegible]
aration may be [illegible]
salt and ammonium [illegible]
ammonium carbonate [illegible]

6. Remove the [illegible]
in a funnel upon a [illegible]
wool. Wash with a [illegible]
wash water gives [illegible]
You now have sodium [illegible]
ration is called the Solvay process.

7. Which is the more [illegible]
the acid carbonate? [illegible]
be converted [illegible]
enough of the [illegible]
normal carbonate and [illegible]
sealed specimen tubes. [illegible]
place in this method of preparing [illegible]
sodium carbonate.

8. **Sodium Chloride, NaCl.** Purification [illegible]
of crude rock salt in [illegible]
diately add calcium [illegible]
0.5 g. lime into a little water [illegible]
solution unt[illegible]
longer for[illegible]
solution [illegible]
ilute h[illegible]
own to [illegible]
rystals [illegible]

gentle heat. Preserve a specimen. Explain use of the calcium hydroxide, of the barium chloride, and of the sodium carbonate, remembering that rock salt usually contains the chloride and sulphate of magnesium.

$$MgCl_2 + Ca(OH)_2 = \text{——},$$
$$MgSO_4 + BaCl_2 = \text{——},$$
$$CaCl_2 + Na_2CO_3 = \text{——},$$
$$\text{(excess) } Na_2CO_3 + HCl = \text{——}.$$

Complete the equations. Common salt may also be purified by passing into a solution of it a current of hydrochloric acid gas.

9. Soap. — Put into an evaporating dish about 50 cc. of a strong solution of sodium hydroxide; add about a cubic centimeter of oil, or tallow. Boil for some time, until the solution becomes thick. Allow the mass to cool, and harden. Describe. What is soap?

10. Flame Test. — Hold a platinum wire in a Bunsen flame until a color is no longer imparted to the flame. Then dip into a solution of common salt and again hold in the flame. Note the color imparted, which is characteristic of sodium compounds.

Potassium

11. Characteristics. — Describe potassium as seen in the class room, noting the same points as in the case of sodium. Compare the two metals in their chemical activity, as seen when exposed to the air, when thrown upon water, etc.

12. Potassium Hydroxide, KOH. — What is the commercial name for this substance? Judging from work with sodium, how do you think it might be prepared?

CaO, and add enough water to reduce to the consistency of cream. Add this milk of lime to the boiling solution of sodium carbonate, and continue to boil several minutes. Allow the precipitate in the solution to settle, and decant the *clear* liquid.

3. **Estimation of the Alkali Prepared.** — Into an evaporating dish, the weight of which is known, measure out half of the clear solution decanted, and boil slowly to dryness. While waiting, proceed with the next work.

4. **Titration of the Alkali.** — Take half of the remaining portion of the decanted liquid, add to it sufficient pure water to make about 30 cc., and titrate half of this with oxalic acid solution. (Refer to work in Chapter X, sections 7 and 8.) Make a second titration, and calculate the total amount of alkali in the half thus used. When the other half has boiled dry, cool and weigh quickly. How do the results by weighing and by titration agree?

5. **Sodium Bicarbonate (Acid-carbonate), $NaHCO_3$.** — (For special students.) Put 50 cc. of strong ammonium hydroxide into a small flask or bottle, and add 15 g. of common salt, finely powdered. Cork the bottle and shake for several minutes, or until the salt is all dissolved. Keep the solution as nearly ice cold as possible. Next, arrange apparatus as shown in Fig. 29; connect with a Kipp generator. Let the carbon dioxide pass through a little water in the wash-bottle, *A*; intro-

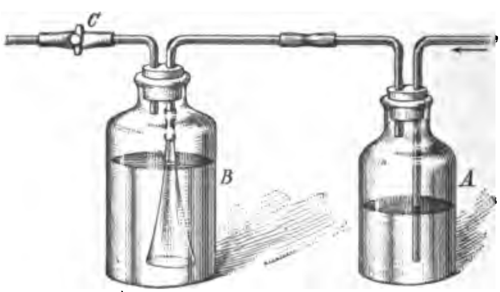

Fig. 29.

duce the gas into the solution in *B* by means of an adapter, that is, a funnel-shaped tube. Why? Except at intervals, keep the clip at *C* closed so that the gas will be under pressure and the absorption more rapid. Keep *B* as cool as possible. Pass the gas for twenty-four hours, or as long as any crystals form. This part of the preparation may be hastened somewhat by putting in with the salt and ammonia solution, at the beginning, as much ammonium carbonate as will dissolve.

6. Remove the crystals from the bottle, and drain them in a funnel upon a platinum cone, or a plug of mineral wool. Wash with a little ice-cold water. Do this till the wash water gives no test for chlorides. (See page 68.) You now have sodium bicarbonate; this method of preparation is called the *Solvay* process.

7. Which is the more soluble in water, the normal or the acid carbonate? How may the acid (bicarbonate) be converted into the normal carbonate? If you have enough of the bicarbonate, convert half of it into the normal carbonate, and preserve the two samples in well-sealed specimen tubes. Write all the reactions that take place in this method of preparing the acid and normal sodium carbonate.

8. Sodium Chloride, NaCl. Purification.—Dissolve 25 g. of crude rock salt in about 85 cc. of hot water, and immediately add calcium hydroxide, made by putting about 0.5 g. lime into a little water. Next, add barium chloride solution until a little of the clear salt solution will no longer form a precipitate with it. After filtering, add a solution of about 1 g. of sodium carbonate, and then dilute hydrochloric acid till no longer alkaline. Boil down to point of crystallization, and cool. Filter out the crystals, wash them with a very little water, and dry with

gentle heat. Preserve a specimen. Explain use of the calcium hydroxide, of the barium chloride, and of the sodium carbonate, remembering that rock salt usually contains the chloride and sulphate of magnesium.

$$\begin{aligned} MgCl_2 + Ca(OH)_2 &= \text{——}, \\ MgSO_4 + BaCl_2 &= \text{——}, \\ CaCl_2 + Na_2CO_3 &= \text{——}, \\ \text{(excess) } Na_2CO_3 + HCl &= \text{——}. \end{aligned}$$

Complete the equations. Common salt may also be purified by passing into a solution of it a current of hydrochloric acid gas.

9. Soap. — Put into an evaporating dish about 50 cc. of a strong solution of sodium hydroxide; add about a cubic centimeter of oil, or tallow. Boil for some time, until the solution becomes thick. Allow the mass to cool, and harden. Describe. What is soap?

10. Flame Test. — Hold a platinum wire in a Bunsen flame until a color is no longer imparted to the flame. Then dip into a solution of common salt and again hold in the flame. Note the color imparted, which is characteristic of sodium compounds.

Potassium

11. Characteristics. — Describe potassium as seen in the class room, noting the same points as in the case of sodium. Compare the two metals in their chemical activity, as seen when exposed to the air, when thrown upon water, etc.

12. Potassium Hydroxide, KOH. — What is the commercial name for this substance? Judging from work with sodium, how do you think it might be prepared?

Write the reaction. What impurities are liable to be in the finished product?

13. Potassium Nitrate, KNO_3. — Describe its appearance and taste. Name two uses for it.

14. Flame Test. — Dip a platinum wire, ignited until perfectly clean, into a solution of some potassium salt, and hold it in the Bunsen flame. Note the color imparted. Now mix with the potassium solution a little sodium chloride, and repeat the flame test. Results? What is your explanation? Observe the flame through a cobalt glass; results? Explain.

CHAPTER XV

ALKALINE EARTH METALS

Compounds of Calcium

1. Lime, Calcium Oxide, CaO. — How is it prepared? What is the chemical reaction?

2. Lime-water, Calcium Hydroxide, $Ca(OH)_2$. — Put a few grams of lime into about 100 cc. of water in a flask, or bottle, and shake for a few minutes. Allow any undissolved portions to settle. Test the clear solution with litmus; results? To which class of oxides does lime belong? (See page 45.) To what class of compounds does lime-water belong? (See "Modern Chemistry," pages 124, 125.)

3. Through 4 or 5 cc. of clear lime-water in a flask blow the breath from the lungs for a moment; results? Continue blowing, or pass a current of carbon dioxide for some time; results? Save this solution,

4. Dissolve a shaving of good soap in a little distilled water. See that the solution is perfectly clear. Take part of the lime-water through which the carbon dioxide was passed in 3, and add a few drops of the soap solution; results? Boil the remainder of the clear solution, saved above, several minutes, then examine it. Has any precipitate formed? Test a portion of the solution just boiled with soap solution. Are the results the same as before boiling? Explain. Remember what caused the precipitate to disappear which formed when you first blew through the lime-water; also the effect of heat upon the solubility of gases. What is such water as this called? Why does a crust form on the inside of boilers and teakettles?

5. Powder finely about $\frac{1}{2}$ g. of calcium sulphate, $CaSO_4$, put it into 200 cc. of pure hot water, and shake for a few minutes to effect solution. Test a few cubic centimeters of this with soap solution; results? Boil another portion and repeat the test; results? What is such water called? Why can you not "soften" it by boiling as you did the other?

6. Hydroxyl Ions. — Remember that hydroxyl is HO. Put 2 or 3 cc. of a solution of copper nitrate or sulphate into each of two tubes. To one add a few drops of sodium hydroxide solution, to the other, a few drops of lime-water. Compare results. What are your conclusions as to the relative character of the ions in each? Repeat, using a solution of ferric nitrate or chloride instead of the copper; results? Write the reactions.

7. Calcium Carbonate, $CaCO_3$. (Uncrystallized.) — To 25 cc. of the spent acid in your Kipp generator for carbon dioxide, add ammonium hydroxide till alkaline, then a solution of ammonium carbonate as long as any precipitate forms. Filter out, wash well with water, and

dry. Describe the product, and preserve it in a specimen tube. Test a portion of it, and prove that it is a carbonate. Write the reaction which represents its preparation.

8. **Flame Test.** — Use a solution of calcium chloride, and repeat the test as for sodium; results?

Compounds of Strontium

9. **Strontium Carbonate, $SrCO_3$.** — Dissolve 5 g. of strontium chloride, $SrCl_2$, or nitrate, in water, and add ammonium carbonate solution as long as any precipitate forms. Filter, wash thoroughly, and dry. Preserve in a specimen tube. Describe and compare with the similar compound of calcium. Write the reaction. From strontium carbonate, how could you prepare the oxide? From the oxide, now prepare the hydroxide? Prepare a small quantity of the oxide, and determine whether it is basic or acidic.

10. **Flame Test.** — Use a solution of strontium chloride or nitrate, and make the flame test; results?

Compounds of Barium

11. **Barium Carbonate, $BaCO_3$.**—Use barium chloride, and prepare the carbonate as you did with strontium. Preserve. Write the reaction.

12. **Barium Oxide.** — Prepare by heating in a small crucible some barium nitrate. Describe its appearance. Determine whether it is an acidic oxide.

13. **Flame Test.** — Make this test in the usual way, and state results.

Compounds of Magnesium

14. Describe metallic magnesium as you have seen it in the form of ribbon. What can you say of the ease with

which it may be oxidized? What forms when it burns in the air?

15. **Magnesium Carbonate, $MgCO_3$.** — Dissolve some magnesium chloride or sulphate in water, and add a solution of ammonium carbonate and a little ammonium hydroxide. Describe the precipitate; write the reaction. Filter, and allow the precipitate to dry. Save for section 16 below. Into each of two tubes put 2 or 3 cc. of a solution of magnesium chloride or sulphate. To one add a gram or more of ammonium chloride, and then to both a solution of ammonium carbonate and a little ammonium hydroxide. How do the results differ? Why? Into one test-tube put a little barium chloride solution; into a second, some strontium chloride solution. To each add some ammonium chloride, and then ammonium carbonate solution. Results? How different from magnesium? Calcium salts behave in this respect like barium and strontium. Suggest a plan for separating magnesium compounds from the others mentioned.

16. **Magnesium Oxide, MgO.** — How have you seen it prepared? Put the magnesium carbonate obtained above into a small crucible, and heat to redness for several minutes. Test a small portion of it with hydrochloric acid; is it still a carbonate? Give the reason for your answer. To another portion add water, and test with litmus paper; result?

CHAPTER XVI

COPPER-SILVER GROUP

Compounds of Copper

1. Cupric Chloride, $CuCl_2$. — Copper forms two classes of salts, — the cuprous and cupric. Cuprous ions are colorless, cupric, blue. Put into a flask 5 g. of copper turnings, nearly cover them with strong hydrochloric acid, and add a little strong nitric acid. Warm gently until the action begins. When the copper is dissolved, decant or filter off from any sediment, and boil down in the hood the clear solution to the point of crystallization. Drain the crystals, and dry them in a desiccator over caustic potash, or even in the air. It may require several days. Preserve them in a tube for future use. Cupric oxide, CuO, instead of copper, may be used to prepare the cupric chloride, by dissolving in strong hydrochloric acid.

2. Cuprous Chloride, Cu_2Cl_2. — Put 5 g. of cupric oxide into a 200 cc. flask, and add 50 cc. of strong hydrochloric acid. Warm gently till the oxide is dissolved; then add 6 or 7 g. of copper turnings, and boil until the solution loses its green color. Pour into a large bottle containing 1 or 2 l. of water. The precipitate is cuprous chloride. Allow it to stand exposed to the light some time; results? Expose a portion of the precipitate to the air; result? Compare in tabular form the two chlorides, in color, form, solubility, effects of light and air.

3. Cupric Sulphate and Nitrate. — When have you prepared these? How? Refer to the work in oxides of nitrogen and sulphur.

4. Cupric Oxide, CuO. Dry Method. — When have you prepared this compound? Put into a small evaporating dish or crucible 5 g. of powdered copper nitrate, and heat gradually to dull redness. What fumes are driven off? Is there any other gas? How could you determine? Preserve your specimen.

5. Wet Method. — Dissolve 5 g. of copper sulphate in 50 cc. of water, and add a solution of potassium or sodium hydroxide as long as any precipitate forms. What is this precipitate? Its color? Boil for a few minutes; what change takes place? Filter out, dry, and preserve. Compare with the cupric oxide made by the previous method.

6. Cuprous Oxide, Cu_2O. — Dissolve 1 g. of sugar in 100 cc. of water, and add 3 or 4 drops of strong hydrochloric acid. Boil 10 to 15 minutes. This will convert the sugar to glucose; neutralize by adding sodium hydroxide solution. Next put into a beaker about 25 cc. of a dilute solution of copper sulphate; add caustic soda solution in excess as in section 5. Now pour in the sugar solution until the blue precipitate dissolves. Heat to boiling. Notice the yellow precipitate of cuprous hydroxide that first forms, and then changes to cuprous oxide. Compare with the cupric oxide. What was the effect of the glucose? What would you have obtained without the sugar?

7. Reactions of the Cupric Ion. — To 2 or 3 cc. of a dilute solution of cupric sulphate add ammonium hydroxide slowly till the precipitate which forms at first redissolves. State results. This is the usual test for a cupric solution. Into each of two test-tubes put 2 or 3 cc. of a cupric sulphate solution. To one add a solution of

potassium cyanide drop by drop till the blue just disappears. Now pass a current of hydrogen sulphide through each. What difference in results? Make a small loop in the end of a platinum wire, heat it and dip it into powdered borax. Hold in the flame again till the borax loses its water of crystallization and becomes perfectly clear. Again heat in the oxidizing flame with a minute particle of cupric oxide. What color is imparted to the bead?

8. Displacement of Cupric Ions. — Put 15 or 20 cc. of a solution of copper sulphate into a beaker, and immerse in it a few bright nails. Allow these to remain till you can no longer see any blue color in the solution. Test a portion of it as in section 7; do you find any indication of cupric ions present? What has become of them? Why do you say so?

Silver

9. Reduction of Silver. — Put into a test-tube about 5 cc. of silver nitrate solution, and to it add dilute hydrochloric acid as long as any precipitate forms. Warm somewhat, and decant the liquid. Put the precipitate into a little cavity in a stick of charcoal, and heat with the reducing flame; results? What other metal have you reduced in this way? This illustrates in a general way the reduction of the more common metals from their ores. (Refer to work in carbon.)

10. Detection of Silver Ions. — The test is simply the reverse of the one used for determining the presence of chlorine ions. Put 1 cc. of a solution of silver nitrate into a test-tube, and add a few drops of hydrochloric acid or of a solution of any chloride; results? Add a little ammonium hydroxide; results? Add nitric acid till the odor of ammonia is not noticeable; results?

11. Interaction of Silver and Halogen Ions. — Refer to work in halogens, in testing chlorine, bromine, and iodine ions. Make a note here of results. Repeat the experiments, if necessary, to refresh your memory.

12. Practical Use of Silver Salts. — Immerse a few sheets of drawing paper or flat-cap in a moderately strong solution of silver nitrate. Dry them in the dark. Next immerse in dilute hydrochloric acid, and again dry in the dark. Instead of immersing the paper, the two solutions may be put on with a brush in a dark room. When dry, put a sheet of this sensitized paper in a printing frame under a negative or leaf and expose to strong sunlight. Notice results.

13. Sepia Prints. — (For special students.) Prepare the following solutions: —

A.	Ferric ammonio citrate	4 g.
	Citric acid	1 g.
	Water	10 cc.
B.	Silver nitrate . . .	1 g.
	Water	4 cc.

When ready for use, mix A and B, and dilute to 20 cc. with water. Apply to white paper with a tuft of cotton, first lengthwise then across the paper, to remove all streaks. Use as little of the solution as possible. Dry in the dark. Print in moderately strong light till shadows are of a golden yellow color. Wash well in water, then immerse in a dilute solution of hypo (1 part in 20) for a minute or two. Again wash thoroughly, and dry. A much more sensitive paper may be made by using oxalic acid and an oxalate instead of the citrate mentioned; but the printing must then be done in light somewhat subdued.

CHAPTER XVII

ZINC, CADMIUM, MERCURY GROUP

1. Solvents for Zinc. — What two acids have you found to react readily with zinc? (Refer to work in hydrogen.) Write the two reactions.

2. The Zinc-copper Couple. — Into each of two test-tubes put a granule of zinc about the same size. Cover one with water, the other with copper sulphate solution. Allow them to stand a few minutes, then add to each 5 or 6 drops of strong sulphuric acid. What difference in action do you notice? Explain.

COMPOUNDS OF ZINC

3. Zinc Sulphate, $ZnSO_4$. — How have you prepared it? Reaction? Describe its appearance by referring to the specimen you preserved. What is the commercial name?

4. Zinc Sulphide, ZnS. — Prepare a solution of zinc sulphate. Put 5 or 6 cc. of it into a test-tube, and add one drop of ammonium sulphide. Note results. Test the solubility of the precipitate in acetic acid; also another portion in sulphuric acid; results?

5. Mass Action. — Dissolve 5 g. zinc sulphate in 150 cc. water. Put 30 cc. of the solution into each of four test-tubes; to the first add one drop of sulphuric acid, to the second about 1 cc. of the same acid, to the third 1 cc. of acetic acid. Through each pass hydrogen sulphide for a few minutes; results? Wherein different? Test the

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fourth with blue litmus paper; does it show an acid reaction? Write the equation for the fourth. Explain the difference between one and two, remembering the only difference in the beginning of the experiment in the case of one, two, and four. How would you write the equation so as to express this *mass action*, showing equilibrium?

6. This action may also be illustrated by dissolving 1 g. each of zinc sulphate and zinc acetate in separate beakers with about 25 cc. of water. Saturate each with hydrogen sulphide. Allow the precipitates to settle, and compare the amounts. If zinc acetate is not to be had, make a solution of zinc sulphate, and add some sodium acetate.

7. Zinc Hydroxide, $Zn(OH)_2$. — Dissolve 10 g. zinc sulphate in 50 cc. water, and add to it slowly a dilute solution of sodium hydroxide as long as the precipitate seems to increase in quantity. Allow it to settle, decant as much of the liquid as possible, and filter. Wash thoroughly. Put a very small portion of the precipitate into each of two test-tubes for further tests. Dry the remainder and preserve. Describe the appearance. To the small portion of the precipitate in one test-tube add a strong solution of sodium hydroxide, and warm gently; results? What must be the character of a substance that will thus react with an alkali? To the zinc hydroxide in the other tube add a little dilute hydrochloric acid; results? What does this reaction indicate regarding the nature of zinc hydroxide?

8. Zinc Oxide, ZnO. — When the hydroxide is dry, put half of it into a small crucible and heat strongly. Describe the oxide thus obtained, both hot and cold. Preserve as a specimen. Put a little of it into water and test with litmus paper; results? What is an important use of zinc oxide?

9. Tests for Zinc Compounds. — The reactions shown with sodium hydroxide, as seen in preparing the hydroxide, and with the sulphide are characteristic, but they must be verified. Put a little of the zinc hydroxide, obtained above, into a small cavity in a piece of charcoal, moisten with a solution of cobaltous nitrate, and heat with the oxidizing flame. A green mass should result. Let the student verify this.

10. Combining Weight of Zinc. — Repeat the zinc-copper experiment above as follows. Put into a beaker 2 g. of freshly powdered copper sulphate, and dissolve it in 25 cc. of water. Immerse in this solution a strip of zinc, weighing 1 or 2 g. (weigh it exactly). Allow it to remain till the blue color has entirely disappeared from the copper solution; then remove the zinc, wash it clean from any adhering particles of copper, dry and weigh it; what is its loss? Determine the amount of copper in the copper sulphate used, remembering to take into account the water of crystallization present. What have you already found to be the combining weight of copper? (See page 57.) From these data find the combining weight of zinc.

At. wt. Cu : At. wt. Zn : : Cu in $CuSO_4$ used : loss of Zn.

Mercury and its Compounds

11. Amalgams. — What is an amalgam? The usual method of preparing? Put into an evaporating dish a few cubic centimeters of a solution of mercurous nitrate and immerse in it a clean penny. After a few minutes remove and rub with a cloth; results? Try in the same way a brass thimble or pin, a bright nail or key, and any other metallic articles available. State which are amalgamated and which are not.

12. Mercurous Nitrate, $HgNO_3$ or $Hg_2(NO_3)_2$. — Put into a small beaker or flask about 5 g. of mercury and add about 10 cc. of dilute nitric acid. Allow it to stand over night. Drain off the spent acid from the crystals, and dry and preserve them for subsequent tests.

13. Mercuric Nitrate, $Hg(NO_3)_2$. — Put into a small flask or beaker about 5 g. of mercury and add about 10 cc. of strong nitric acid, warm gently till the mercury is all dissolved. Transfer to an evaporating dish, and boil a few minutes. To be sure the action is complete, put a few drops of the solution into a test-tube, dilute with water, and add a few drops of sodium chloride solution. If a white precipitate forms upon adding the salt solution, add a little more nitric acid and boil again. Evaporate till a drop of the solution cooled on a glass rod will crystallize. Allow the whole to cool, drain off, and dry the crystals. Preserve them for further tests.

14. Distinctions between Mercuric and Mercurous Ions. — Dissolve, separately, portions of the two salts prepared above, in 10 to 15 cc. of water, to which you have added 2 or 3 drops of strong nitric acid. Make four tests with each solution, and tabulate the results. *a.* To a portion of each in separate test-tubes add slowly, drop by drop, a little potassium iodide solution; results? *b.* To another portion of each add dilute hydrochloric acid; results? *c.* To a third portion of each add ammonium hydroxide; results? *d.* In the same way, try a solution of stannous chloride, adding slowly.

15. Unknown Solutions. — Let the teacher furnish two or three unknown solutions, and have the student determine by some of the above tests to which class of mercury salts each belongs; record results. What tests do you prefer?

16. Mercuric Oxide, Yellow. — Dissolve 5 g. of powdered mercuric chloride, $HgCl_2$, in 100 cc. of hot water, and add potassium hydroxide solution as long as any precipitate forms. Filter out the bright yellow precipitate, and wash several times with a considerable quantity of water. Dry and preserve in a specimen tube. What other oxide have you prepared in a similar way? What compound have you generally obtained upon adding potassium hydroxide? How do you account for the fact that the oxide is formed?

17. Dry Test for Mercury. — Mix a few crystals of some salt of mercury, such as the nitrate, with a little dry sodium carbonate, put into a hard glass tube and heat strongly. Notice the deposit on the sides of the tube; what is it?

CHAPTER XVIII

ALUMINUM AND ITS COMPOUNDS

1. Characteristics of Aluminum. — Examine a piece of the metal and describe it as to color, density, hardness, melting-point, and ease of oxidation. Test it in the last two respects by heating in the oxidizing flame. What is the effect upon aluminum of strong hydrochloric acid? Test its solubility in strong nitric and sulphuric acids, both cold and heated; results? What is the character of a substance that will dissolve in acids? What gaseous product was given off in each case? Determine the effect of a strong solution of sodium hydroxide; results? What gas is formed? What is the nature of a substance that will dissolve in alkalies? How do you account for the character of aluminum shown by the above tests with acids and alkalies?

Compounds of Aluminum

2. **Alums.** — Dissolve in 50 cc. of warm water 10 g. of common alum, powdered, and suspend a thread in the solution. Allow it to cool slowly. What is the shape of the crystals formed? Make a drawing to represent one.

3. **Preparation of Common Alum.** — Dissolve 10 g. of powdered aluminum sulphate in 25 cc. of hot water, and in another beaker 2.6 g. of potassium sulphate in 10 cc. of hot water. If the aluminum sulphate used is anhydrous, take 5 g. of potassium sulphate instead of 2.6, and dissolve it in 20 cc. of water. Pour the two solutions together, stir well, and let the compound cool. Crystallize on a thread, as before. Examine the crystals. Determine whether they contain water of crystallization. What kind of a salt is alum? Preserve your specimen.

4. **Aluminum Hydroxide, $Al_2(OH)_6$.** — What method would you suggest for preparing this compound? Put a few cubic centimeters of a solution of alum into a test-tube, and attempt to prepare the hydroxide; results?

5. **Hydrolysis of Aluminum Compounds.** — Put into a test-tube 5 cc. of a moderately strong solution of alum, or aluminum sulphate; fit to the tube a cork with delivery tube. Have at hand a test-tube containing clear lime-water. Now add to the solution of alum some sodium carbonate solution; quickly insert the cork, and conduct the gas into the lime-water; results? What is the gas? When you added a carbonate to solutions of barium, strontium, etc., what kind of a precipitate did you obtain? Was any gas evolved?

6. Determine whether the precipitate now is a carbonate. Wash it by decantation several times with considerable quantities of water. Transfer it to a test-tube, and

add a little hydrochloric acid. Is any gas evolved, indicated by effervescence, as is always the case with a carbonate?

7. Try in the same way to prepare the sulphide by adding a few drops of ammonium sulphide to an alum solution; notice the odor of the gas evolved when you add the ammonium sulphide. Filter out the precipitate, wash thoroughly with water, and then test it for a sulphide according to method given on page 68. What are your conclusions? What do you mean by hydrolysis?

8. Tests for Aluminum in Compounds.—Add ammonium sulphide in small quantity to the solution to be tested. State results as obtained in preceding section, and give color and appearance of the precipitate.

9. Add sodium hydroxide solution drop by drop and then in excess to the solution to be tested. At the close warm gently, if necessary; results? Compare with a similar test for zinc ions.

10. Make the charcoal test with aluminum hydroxide as you did with zinc, and state wherein the results are different in this case. Sometimes this test is made by folding the filter paper containing the zinc or aluminum hydroxide and wrapping a platinum wire around it. Moisten the paper and precipitate with a solution of cobaltous nitrate, and hold in the Bunsen flame. Results as already described.

CHAPTER XIX

TIN AND LEAD

1. **Preparation of the Chlorides of Tin.** — Tin forms two classes of salts, — stannous and stannic. Put about 1 g. of granulated tin or tin foil into a beaker, and add 10 cc. of strong hydrochloric acid. Warm gently. When most of the tin is dissolved, decant the clear solution and evaporate to dryness on a water bath. This should give you stannous chloride, $SnCl_2$. Describe its appearance. Save for future tests.

2. Put about 1 g. of tin into a beaker or flask, add 10 cc. of strong hydrochloric acid and about 3 cc. of strong nitric acid. Warm gently. When most of the tin is dissolved, saturate it with chlorine. Decant the clear solution, and evaporate to dryness. Describe. You should have prepared thus stannic chloride, $SnCl_4$. Save for tests in next section.

3. **Distinctions between Stannous and Stannic Salts.** — Put a little of a solution of each class of salts prepared above into separate test-tubes; add a drop or two of hydrochloric acid if necessary to secure a clear solution. To each add slowly a solution of mercuric chloride, $HgCl_2$. Note results in tabular form. To a fresh portion of each, acidulated with hydrochloric acid, add hydrogen sulphide; results? Test the solubility of these precipitates in ammonium sulphide. Warm, if necessary; results?

4. **Stannic Oxide, SnO_2.** — Put 2 or 3 g. of tin into a test-tube and cover with strong nitric acid. Result?

When the tin is all decomposed, fill the tube with water and filter. Dry the precipitate. What is it? When dry, heat it strongly. Give appearance hot and cold. Compare with zinc compounds heated thus.

Lead and its Compounds

5. **Reduction from its Ores.** — How have you already obtained metallic lead from its compounds? Refer to work in carbon. Write the reaction.

6. **Characteristics of Lead.** — From observation note the color, hardness, and malleability of lead. Determine whether easily melted; also whether easily oxidized.

7. **Lead Nitrate, $Pb(NO_3)_2$.** — Dissolve 5 g. of lead, cut into small pieces, in nitric acid; decant from any sediment and evaporate to crystallization. Filter out and dry the crystals. Preserve the crystals in a specimen tube. Describe them, and write the reaction, noting what gas appeared. Refer to preparation of nitric oxide. Lead oxide, PbO, may be used instead of metallic lead, if necessary.

8. **Lead Chloride, $PbCl_2$.** — Dissolve a little lead nitrate in water, and to the solution add dilute hydrochloric acid; results? Let the precipitate settle, and decant the supernatant liquid. Transfer the precipitate to a test-tube, add 10 or 15 cc. of water, and heat to the boiling-point; results? Cool rapidly; results? Compare the product you have now with what you had before heating.

9. **Lead Sulphate, $PbSO_4$.** — Dissolve 5 g. of lead nitrate in 50 cc. of water, add a little alcohol, and then dilute sulphuric acid as long as any precipitate forms. Filter, and wash the precipitate well. Dry and preserve

the specimen. Describe it. Write the reaction which took place. What was the use of the alcohol?

10. Lead Chromate, Chrome Yellow, $PbCrO_4$. — Dissolve 5 g. of lead nitrate in 50 cc. of water, and add a solution of potassium dichromate as long as any precipitate forms. Filter out, wash thoroughly, and dry. Preserve in a specimen tube. Describe the compound. Write the reaction.

11. White Lead, Basic Lead Carbonate. — Dissolve 5 g. lead nitrate in 50 cc. of water, and add a solution of sodium carbonate till precipitation is complete. Wash and dry the precipitate. Preserve in a specimen tube. Describe its appearance. What use has white lead?

12. Oxides of Lead. — What is the source of the litharge of commerce? What is its formula?

13. Lead Dioxide, PbO_2. — (Special students.) Do the work in the hood. Dissolve 10 g. of lead acetate in 40 cc. of water and add 40 cc. of a solution of caustic soda, containing about 8 or 9 g. of solid sodium hydroxide. Notice the results. What is the precipitate? Write the reaction. Next pass in chlorine until the oxidation is complete. Filter or decant, and wash the precipitate. Transfer to a beaker and boil with 50 to 60 cc. of very dilute nitric acid. Decant and wash two or three times. Dry the precipitate, and describe. Preserve a specimen. What purpose does the chlorine serve? Reaction?

14. Means of detecting Lead Salts. — The usual tests for lead in solution you have already made. Note here the reaction of lead ions with *a*. hydrochloric acid; *b*. potassium dichromate; *c*. sulphuric acid. Heat strongly some lead compound, as the nitrate, in a hard glass tube; results? Compare with zinc and tin.

15. Displacement of Lead Ions by Zinc. — Dissolve about 3 g. of lead acetate in 50 cc. of water and immerse in it a

strip of zinc weighing 2 or 3 g. Weigh both the lead salt and the zinc accurately. Allow the solution to stand over night. Remove the zinc from the solution, carefully wash off all adhering particles of lead, dry, and again weigh. Find the loss of the zinc. Determine the amount of lead in the lead acetate used, and from this calculate the combining weight of lead, using for zinc that determined on page 83.

CHAPTER XX

ARSENIC, ANTIMONY, BISMUTH

1. **Characteristics of Arsenic.** — Describe metallic arsenic in general appearance. Test its malleability. Put a crystal of it into a crucible or small evaporating dish, and heat gently. Notice the odor. Does the crystal melt?

Compounds of Arsenic

2. **Arsine, Hydrogen Arsenide, AsH_3.** — This is a very poisonous gas and must be prepared with care. The experiment is better suited to the lecture room, unless the classes are very small. Prepare a flask for the generation of hydrogen and attach a jet. When all the air has been expelled (see work in hydrogen), light the hydrogen. Hold a cold dish against the flame. No spot should be left upon the dish, if the zinc is pure.

3. Now pour into the generating flask a few drops of an arsenic solution prepared as in 4 below. Again light the jet. Note the change in the appearance of the flame. What else do you notice different from what you observed in the case of the burning hydrogen? Hold a beaker over the flame, but not touching it; what two deposits do

you see forming upon the beaker? Hold a cold dish against the flame; what is the deposit? Describe its appearance. Test its solubility in nitric acid; in a solution of bleaching powder; results? Why did the cold dish bring about this deposit? Reactions.

4. **Arsenic Trioxide, As_2O_3.** — What is its general appearance? Test its solubility in strong hydrochloric acid. Use only a very little of the trioxide on the tip of a knife blade with 2 or 3 cc. of acid; warm gently. What kind of oxides would you expect to be soluble in acids? (Refer to page 45.) Save this solution for 5 below. Test the solubility of arsenic trioxide in a strong solution of sodium hydroxide. Use a very little of the oxide as in the first test, and warm it; results? What kind of oxides, as a rule, are soluble in alkalies? What are your conclusions as to the nature of arsenic trioxide? What compound does it form with hydrochloric acid? With sodium hydroxide? What other elements have you found showing thus a double character? Reactions.

5. **Arsenic Trisulphide, As_2S_3.** — Through the solution of arsenic trioxide in hydrochloric acid, saved above, pass a current of hydrogen sulphide; results? Test the solution with litmus paper; if acid, neutralize with ammonium hydroxide, then add ammonium sulphide, and warm gently; results?

6. **Tests for Arsenic.** — The most delicate test, and the one commonly used when arsenic is present in small quantities only, is what is known as Marsh's test. This is the one described in the preparation of arsine.

7. **Scheele's Test.** — Put 2 or 3 cc. of a solution of copper sulphate into a test-tube, and add ammonium hydroxide, drop by drop, until the precipitate which forms at first just dissolves. Then add the supposed arsenic solution

slowly. This is an easy test when considerable quantities of arsenic are present. A bright green precipitate will form.

8. **Dry Test.** — Put into a hard glass tube sealed at one end a very little of an arsenic compound, such as the trioxide, mixed with a little powdered charcoal. Just above the mixture put a plug of charcoal, snugly fitting. Heat the charcoal red hot and then heat the mixture strongly; results? Compare with the arsine spots. Compare with mercury test.

Antimony and its Compounds

9. **Characteristics.** — If this metal is not to be had in the laboratory, it may easily be reduced from a compound on charcoal. Mix a little antimony tartrate with powdered sodium carbonate, put into a cavity on charcoal, and heat strongly with the reducing flame. Make a note of the physical characteristics usually observed.

10. **Antimony Trichloride, $SbCl_3$.** — It may be prepared by dissolving the metal in aqua regia and evaporating.

11. **Hydrolysis of Antimony Trichloride and Mass Action.** — Put a crystal of antimony trichloride into a test-tube, and add 4 or 5 cc. of water; results? Test the solution with blue litmus; results? Explain.

$$SbCl_3 + H_2O = SbOCl + 2\,HCl.$$

Now warm the solution and add slowly, a drop at a time, strong hydrochloric acid. Add no more than is necessary to dissolve the precipitate. Now pour this solution into a beaker nearly full of water; results? Put a portion of the new solution into a test-tube, warm, and again add hydrochloric acid; results? We may write the reaction, thus: —

$$SbOCl + 2\,HCl = SbCl_3 + H_2O.$$

This is what we call a reversible reaction, and illustrates *mass action*. When the water is in excess, we have the decomposition of the antimony trichloride as shown by the first reaction; when the quantity of the hydrochloric acid is increased, the oxychloride, SbOCl, is again dissolved, forming the trichloride. We would then read the equation in the reverse order. Compare effect of hydrogen sulphide upon zinc compounds.

12. Is the antimony trichloride completely hydrolyzed? Put a few crystals of the compound into 5 cc. of water; let the white precipitate settle, and filter or decant the clear liquid. Through this clear solution pass a current of hydrogen sulphide. Does a precipitate form showing some antimony in solution? To this add ammonium sulphide and warm gently; results? Compare with arsenic trisulphide treated the same way. Describe the antimony trisulphide prepared.

13. Stibine, Antimoniureted Hydrogen, SbH_3. — This may be prepared as arsine was, using a little antimony trichloride solution. Compare the spots with those of arsenic. Try the effects of nitric acid and of a solution of bleaching powder; results? Wherein are they different from those observed when arsenic was used?

Bismuth and its Compounds

14. Reduction. — Put a little bismuth nitrate, $Bi(NO_3)_3$, mixed with sodium carbonate, into a small cavity in charcoal and heat strongly with the reducing flame. Make a note of the usual metallic characteristics. Test the solubility of bismuth in nitric acid; results?

15. Bismuth Nitrate, $Bi(NO_3)_3$. — Put a few crystals of bismuth nitrate into about 5 cc. of water in a test-tube;

results? Test with litmus; results? Warm and add nitric acid, a drop at a time. Be careful not to add an excess; result? Pour into a large beaker of water; result? What other compound have we seen hydrolyzed the same way? Write an equation showing the reversible action in this case and explain.

16. Bismuth Sulphide, Bi_2S_3. — Again put a few crystals of bismuth nitrate into water, warm gently, and decant or filter off the clear solution. Pass a current of hydrogen sulphide through the filtrate; result? What does this show regarding the extent of the hydrolysis? Decant as much of the solution as possible. Add some ammonium sulphide to the black precipitate and warm gently; result? How different from the action upon arsenic and antimony sulphide? Suppose you had a solution of arsenic, antimony, and bismuth chlorides, how could you separate them?

CHAPTER XXI

IRON, NICKEL, AND COBALT

1. Compounds of Iron. — Iron forms two classes of salts, —ferrous and ferric. Let the student distinguish between them. Arrange results in tabular form. Take as your examples ferrous sulphate and ferric chloride. Test separate portions of each with ammonium hydroxide, potassium sulphocyanide, and potassium ferrocyanide. Of the last use only a few drops.

2. Ferrous Salts, Unstable. Proof. — Powder two or three small crystals of ferrous sulphate, and dissolve in cold water. Set half of the solution aside. To the other half add a few drops of nitric acid, and boil a few minutes.

What change in color occurs? Test this solution further with potassium sulphocyanide; what chemical change has taken place? After making the next experiment test the first half of the original solution to determine whether it has changed.

3. Reduction of Ferric Salts. — Put about 10 cc. of a solution of ferric chloride into a large test-tube, drop into it a large granule of zinc, and add strong hydrochloric acid till a brisk effervescence begins; what is the gas? What kind of an agent is it in the nascent form? Mention some other like agent and some work done with it. In 5 or 6 minutes a decided change should be noticeable. How does the solution appear now? Test it to learn what chemical change has occurred; conclusion? Write the reaction.

4. Other Iron Ions. — *a.* Put 5 or 6 cc. of a fresh solution of potassium ferrocyanide, K_4FeCy_6, into a test-tube, and add a little potassium sulphocyanide solution; have you any evidence of the presence of ferric ions? Try another portion with ammonium hydroxide; do you obtain a test for either ferrous or ferric ions? Conclusion? *b.* In the same way try a fresh solution of potassium ferricyanide, K_3FeCy_6. Conclusion? Would it be correct to regard the reagents just tested as double cyanides, as alum is a double sulphate? (Refer to alum.) Could we write the formulæ K_4Cy_4, $FeCy_2$ instead of K_4FeCy_6, and K_3Cy_3, $FeCy_3$ instead of K_3FeCy_6? Give reason for your answer. If so, what ions would there be in a solution of each compound?

5. Preparation of Ferrous Sulphate, $FeSO_4$. — Put into a 200 cc. flask 10 g. of iron turnings or fine wire, and add 50 to 75 cc. of dilute sulphuric acid. Allow the action to continue without heating until the iron is all dissolved.

If necessary, let it remain over night. When ready to proceed, test a small portion to determine whether it contains ferric ions. If so, add a few more filings and acid, if necessary, to start the action afresh. The ferric ions will soon be reduced to the ferrous; why? Add considerable water to dissolve any crystals that may have formed, and decant or filter the solution from any sediment. Boil it down in an evaporating dish as rapidly as possible to point of crystallization; why? Allow it to cool, drain off the mother liquor from the crystals, and dry them between blotting paper. Quickly seal in tightly stoppered bottles; why?

6. Preparation of Ferric Chloride, $FeCl_3$. — Suggest a plan for preparing ferrous chloride and then converting it into ferric chloride. Name some other oxidizing agent besides nitric acid.

7. Preparation of Ferric Alum. — (For special students.) What is an alum? How have you prepared common alum? The plan is the same here. Powder 5 g. of fresh ferrous sulphate and dissolve in 8 or 10 cc. of water to which has been added 1 g. of strong sulphuric acid. Heat to boiling and add a few drops of nitric acid. Boil a few minutes till the solution is thoroughly oxidized. Add a little more acid if necessary to do this. Next dissolve about 1 g. of ammonium sulphate in *very* little boiling hot water, and add it to the solution of ferric sulphate. Allow it to cool. If crystals do not form, concentrate a little further and again cool. Describe the crystals and compare them with those of common alum. What is the formula for the alum just prepared?

8. Nickel Compounds. — (The salts of nickel are unimportant and this work is at the option of the instructor.) Examine any nickel compounds at hand; note the color

and general appearance. Make a borax bead in the usual way, and fuse with it in the oxidizing flame a ***very small*** particle of some compound of nickel. If the bead is black, it is because you have used too much of the compound. Repeat the experiment if necessary; results?

9. Reactions. — Put 4 or 5 cc. of a solution of any nickel salt into a test-tube and add a few drops of ammonium sulphide; result? Write the equation. Test the solubility of the precipitate in nitric acid; result? To another portion of a nickel solution add a little sodium or potassium hydroxide solution, and then in excess; results? Write the equation.

10. Compounds of Cobalt. — (Work unimportant.) Make the bead test as for nickel, and compare results. Make a solution of nickel and cobalt in about equal proportions, and with this repeat the bead test. Which one of them can you not detect now?

11. Reactions. — Try ammonium sulphide as with nickel. Compare results. Write the equation. Try also sodium or potassium hydroxide. Compare. Write equation.

12. Sympathetic Inks. — With a match as a pen write with a solution of cobalt nitrate upon paper. Let it dry, and notice that the letters are scarcely visible. Now heat the paper; result? Explain.

CHAPTER XXII

CHROMIUM AND MANGANESE

1. Compounds of Chromium.* — Chromium forms a variety of compounds, in some respects similar to iron, in others more nearly like those of manganese.

2. Chromic Ions. — To a few cubic centimeters of any chromic solution, as chromic chloride or sulphate, add cautiously a few drops of ammonium or sodium hydroxide; result? What is the compound formed? Write the reaction.

3. Other Chromium Salts. — Chromous compounds are even more unstable than ferrous compounds and will not be found in many laboratories. Test a solution of potassium chromate, K_2CrO_4, and one of potassium dichromate, $K_2Cr_2O_7$, with some alkali, as above. Do you obtain any evidence of the presence of chromic ions? What salts of iron do these compounds resemble in this particular?

4. Conversion of Chromic Salts into Chromates. — This is accomplished by oxidation. Mix together about 1 g. of potassium nitrate and 1 g. of sodium carbonate, and put with the mixture some chromic hydroxide or some chromic salt. Put the mixture on platinum foil or on a piece of broken evaporating dish, and heat to redness until the mass is thoroughly melted. Allow it to cool,

* If no chromic salts are at hand, the instructor may easily prepare them by any of the methods of reduction suggested below.

and dissolve it in water. Judging from the color, what compound of chromium do you think you have obtained? Potassium nitrate is an oxidizing agent.

5. Oxidation of Chromates. — Take two test-tubes of the same size. Into one put 5 cc. of a moderately strong solution of potassium chromate; into the other 5 cc. of potassium dichromate solution; compare the colors of the two. To the first add a few drops of strong nitric acid; result? Compare the colors again; conclusion? Reaction? What kind of an agent is nitric acid? State its effect upon a ferrous solution; upon tin.

6. Summary. — Being given a chromous compound, state how you would change it, first, to a chromic compound; second, to a chromate; third, to a dichromate.

7. Reduction of Potassium Dichromate. — Repeat section 5 above, except that you may add a little strong potassium hydroxide solution to the potassium dichromate instead of acid to the other. Compare colors before and afterward. What chemical change has occurred? Equation?

8. Reduction of Chromates. — Almost any of the reducing agents we have studied may be used. Put into a test-tube or small flask 5 or 10 cc. of potassium chromate (or dichromate) solution, and introduce into it a current of sulphur dioxide. Refer to method of preparing it. What change in color occurs? What does this indicate? Test for chromic ions; result? Conclusion? Try also hydrogen sulphide, passing it through an acidulated solution of potassium dichromate. What two results? Study out the reaction and write the equation. Name two other reducing agents you have used previously.

9. Summary. — Being given a solution of potassium dichromate, state how you might convert it into a chromate, into a chromic salt, and into a chromous salt.

10. Some Insoluble Chromates. — *a. Lead chromate.* How have you already prepared it? *b. Silver chromate,* Ag_2CrO_4. Put 4 or 5 cc. of silver nitrate solution into a test-tube, and add a little potassium chromate solution. Describe the precipitate. Write the equation. *c. Barium chromate,* $BaCrO_4$. To a little of a solution of barium chloride in a test-tube add a few drops of a potassium dichromate solution; result? Compare with lead chromate. Write the equation.

11. Chromic Sesquioxide, Cr_2O_3. — Make the bead test, using any chromium solution; result? The color is due to the chromic sesquioxide formed. This oxide is basic.

12. Chromic Anhydride, CrO_3. — Of what acid is this compound theoretically the anhydride? Can you prepare the acid by adding water? *Preparation.* (For special students.) Dissolve 10 g. powdered potassium dichromate in about 15 cc. of hot water; add slowly 14 cc. of strong sulphuric acid, with constant stirring. Allow the solution to stand over night for the potassium acid sulphate to crystallize out. Decant the red solution from the crystals. Heat nearly to boiling, and add about 5 cc. more of strong sulphuric acid. Boil slowly until crystals begin to form on the surface of the liquid. Allow it to stand over night, and decant the liquid from the crystals. Wash them with a very little strong nitric acid, and dry by gentle heating. As soon as dry, put them into a warm glass tube and seal hermetically. Boil down the liquid decanted from these crystals till new ones form. Put a few of them upon a piece of filter paper, and notice what happens; explain. Study the reactions that have occurred in the above work.

13. Practical Work. — What did you find to be the effect of nitric acid upon a ferrous solution? Potassium

dichromate is also a strong oxidizing agent. Try one of the following experiments. *a.* Put into a test-tube 3 or 4 cc. of a solution of ferrous sulphate, and acidify with hydrochloric acid. Now add, drop by drop, a solution of potassium dichromate until you can see no further change; what visible result? Take a part of it, dilute somewhat with water, and test for ferric ions with potassium sulphocyanide; result? Conclusion?

14. *b.* Powder a little ferrous sulphate, and quickly weigh out 1 g. of it. Dissolve this in 50 cc. of water, and add about 5 cc. dilute sulphuric acid. Prepare a potassium dichromate solution by dissolving .49 g. of the crystals in 100 cc. of water. Prepare also a few cubic centimeters of a solution of potassium ferricyanide. This, with a ferrous salt, gives a blue precipitate; with a ferric, a brown coloration, and is to be used to show when the iron is completely oxidized. Fill a burette with the dichromate solution, and take the reading. In three or four places upon a dry porcelain dish put a drop of the ferricyanide solution.

15. Put 10 cc. of the ferrous sulphate solution into a flask, and run in the potassium dichromate slowly until a drop of the solution on a stirring rod ceases to turn a drop of the ferricyanide blue, but changes it to brown. Again read the burette. Make two or three titrations thus. Each cubic centimeter of dichromate used corresponds to .0056 g. of iron. Calculate the amount of iron in the 1 g. of ferrous sulphate used, remembering the water of crystallization, seven molecules. How does this agree with the amount found by titration? Study out the chemical changes that have taken place, and write the reactions.

16. Compounds of Manganese. — In many respects they resemble those of chromium. The manganous salts and the permanganates are the only common ones.

17. Manganous Reactions. — To a few cubic centimeters of a solution of manganous chloride, $MnCl_2$, in a test-tube, add a little potassium hydroxide solution. Note the color of the precipitate. What is it? Is it soluble in excess of potassium hydroxide? In the same way try the effects of ammonium sulphide upon a manganous solution; result? Determine whether the precipitate is a sulphide. To do this, filter out the precipitate and wash it well, then add to it hydrochloric acid; result? Conclusion? Write the two reactions. Make the bead test with some manganous salt, as you did with nickel and cobalt; result?

18. Permanganate Ions. —Make a very dilute solution of potassium permanganate, $KMnO_4$, and test it with potassium hydroxide solution. Do you obtain any evidence of the presence of manganous ions? Give reason for your answer.

19. Practical Uses of Potassium Permanganate. — Like potassium dichromate, the permanganate is a strong oxidizing agent. The following experiments will illustrate this: Put into a beaker 50 cc. of some water which has been discolored by organic matter, as cistern water from a new shingled roof, and add a little strong sulphuric acid. Heat upon a wire gauze nearly to boiling, and add a solution of potassium permanganate slowly until the water *begins* to turn pink. What change has taken place in the water? What reactions take place? Write the equations.

20. Try either of the following: *a.* Put 5 cc. of a dilute solution of ferrous sulphate into a 100 cc. flask and add 1 or 2 cc. of sulphuric acid. Without heating, gradually add a solution of potassium permanganate until it begins to show pink. Test for ferric ions by potassium sulphocyanide or ammonium hydroxide; conclusion? Explanation? By what other reagent have you thus changed a

17. Manganous Reactions. — To a few cubic centimeters of a solution of manganous chloride, $MnCl_2$, in a test-tube, add a little potassium hydroxide solution. Note the color of the precipitate. What is it? Is it soluble in excess of potassium hydroxide? In the same way try the effects of ammonium sulphide upon a manganous solution; result? Determine whether the precipitate is a sulphide. To do this, filter out the precipitate and wash it well, then add to it hydrochloric acid; result? Conclusion? Write the two reactions. Make the bead test with some manganous salt, as you did with nickel and cobalt; result?

18. Permanganate Ions. —Make a very dilute solution of potassium permanganate, $KMnO_4$, and test it with potassium hydroxide solution. Do you obtain any evidence of the presence of manganous ions? Give reason for your answer.

19. Practical Uses of Potassium Permanganate. — Like potassium dichromate, the permanganate is a strong oxidizing agent. The following experiments will illustrate this: Put into a beaker 50 cc. of some water which has been discolored by organic matter, as cistern water from a new shingled roof, and add a little strong sulphuric acid. Heat upon a wire gauze nearly to boiling, and add a solution of potassium permanganate slowly until the water *begins* to turn pink. What change has taken place in the water? What reactions take place? Write the equations.

20. Try either of the following: *a.* Put 5 cc. of a dilute solution of ferrous sulphate into a 100 cc. flask and add 1 or 2 cc. of sulphuric acid. Without heating, gradually add a solution of potassium permanganate until it begins to show pink. Test for ferric ions by potassium sulphocyanide or ammonium hydroxide; conclusion? Explanation? By what other reagent have you thus changed a

17. Manganous Reactions. — To a few cubic centimeters of a solution of manganous chloride, $MnCl_2$, in a test-tube, add a little potassium hydroxide solution. Note the color of the precipitate. What is it? Is it soluble in excess of potassium hydroxide? In the same way try the effects of ammonium sulphide upon a manganous solution; result? Determine whether the precipitate is a sulphide. To do this, filter out the precipitate and wash it well, then add to it hydrochloric acid; result? Conclusion? Write the two reactions. Make the bead test with some manganous salt, as you did with nickel and cobalt; result?

18. Permanganate Ions. —Make a very dilute solution of potassium permanganate, $KMnO_4$, and test it with potassium hydroxide solution. Do you obtain any evidence of the presence of manganous ions? Give reason for your answer.

19. Practical Uses of Potassium Permanganate. — Like potassium dichromate, the permanganate is a strong oxidizing agent. The following experiments will illustrate this: Put into a beaker 50 cc. of some water which has been discolored by organic matter, as cistern water from a new shingled roof, and add a little strong sulphuric acid. Heat upon a wire gauze nearly to boiling, and add a solution of potassium permanganate slowly until the water *begins* to turn pink. What change has taken place in the water? What reactions take place? Write the equations.

20. Try either of the following: *a.* Put 5 cc. of a dilute solution of ferrous sulphate into a 100 cc. flask and add 1 or 2 cc. of sulphuric acid. Without heating, gradually add a solution of potassium permanganate until it begins to show pink. Test for ferric ions by potassium sulphocyanide or ammonium hydroxide; conclusion? Explanation? By what other reagent have you thus changed a

17. Manganous Reactions. — To a few cubic centimeters of a solution of manganous chloride, $MnCl_2$, in a test-tube, add a little potassium hydroxide solution. Note the color of the precipitate. What is it? Is it soluble in excess of potassium hydroxide? In the same way try the effects of ammonium sulphide upon a manganous solution; result? Determine whether the precipitate is a sulphide. To do this, filter out the precipitate and wash it well, then add to it hydrochloric acid; result? Conclusion? Write the two reactions. Make the bead test with some manganous salt, as you did with nickel and cobalt; result?

18. Permanganate Ions. —Make a very dilute solution of potassium permanganate, $KMnO_4$, and test it with potassium hydroxide solution. Do you obtain any evidence of the presence of manganous ions? Give reason for your answer.

19. Practical Uses of Potassium Permanganate. — Like potassium dichromate, the permanganate is a strong oxidizing agent. The following experiments will illustrate this: Put into a beaker 50 cc. of some water which has been discolored by organic matter, as cistern water from a new shingled roof, and add a little strong sulphuric acid. Heat upon a wire gauze nearly to boiling, and add a solution of potassium permanganate slowly until the water *begins* to turn pink. What change has taken place in the water? What reactions take place? Write the equations.

20. Try either of the following: *a*. Put 5 cc. of a dilute solution of ferrous sulphate into a 100 cc. flask and add 1 or 2 cc. of sulphuric acid. Without heating, gradually add a solution of potassium permanganate until it begins to show pink. Test for ferric ions by potassium sulphocyanide or ammonium hydroxide; conclusion? Explanation? By what other reagent have you thus changed a

ferrous solution? *b. Estimation of iron.* Prepare a standard solution of potassium permanganate by dissolving .316 g. of the crystals in 100 cc. of distilled water, and with it fill a burette. Without heating, dissolve 1 g. of ferrous sulphate, freshly powdered, in 50 cc. of water, and add 5 cc. of sulphuric acid. Take 10 cc. of this solution in a small flask, and after reading the burette, run in the permanganate solution somewhat rapidly at first, more slowly afterward, until, after shaking, the solution remains pink. Again read the burette and make a note of the amount of potassium permanganate solution used in oxidizing the iron. Calculate the amount of iron present. The solution of permanganate as prepared is such that 1 cc. is equivalent to .0056 g. of iron. How does your result agree with the actual amount of iron in 1 g. of pure ferrous sulphate? It is customary to make two or three titrations so as to verify your work.

APPENDIX

CHEMICALS FOR CLASS OF TEN

Acetic acid	1 lb.
Alcohol	1 pt.
Alum	2 lb.
Aluminum sulphate	¼ "
Ammonium carbonate	¼ "
" chloride	1 "
" ferric citrate	½ oz.
" hydroxide	4 lb.
" nitrate	½ "
" sulphate	⅛ "
" sulphide	1 "
Antimony, metallic	¼ "
" potassium tartrate	¼ "
" trichloride	½ oz.
Arsenic, metallic	¼ "
" trioxide	⅛ lb.
Barium chloride	¼ "
" nitrate	¼ "
" oxide	⅛ "
Bark charcoal	1 "
Bismuth, metallic	¼ "
" nitrate	¼ "
Bleaching powder	¼ "
Borax	½ "
Bromine	1 oz.
Calcium carbide	½ lb.
" chloride	⅛ "
" fluoride	⅛ "
" sulphate	⅛ "
Candles	
Carbon disulphide	¼ "
Carmine	¼ oz.
Charcoal, animal, powdered	½ lb.
" stick	1 doz.
" wood, powdered	½ lb.
Chromium nitrate	1 oz.
Citric acid	½ "
Cobalt nitrate	⅛ lb.
Cochineal	¼ oz.
Copper nitrate	⅛ lb.
" oxide	½ "
" sulphate	¼ "
" turnings	½ "
Ether	¼ "
Filter papers, 5 in.	500
Fuming nitric	¼ lb.
Fusel oil	¼ "
Gunpowder	½ "
Hydrochloric acid	10 "
Indigo	½ oz.
Iodine	½ "
Iron chloride (ferric)	½ lb.
" filings	½ "
" sulphate	½ "
" sulphide	2 "
" wire	¼ "
Lead acetate	⅛ "
" metallic	½ "
" nitrate	¼ "
" oxide, litharge	¼ "
" " minium	¼ "
Lime	1 "

Litmus	½ oz.
" papers, red and blue	1 q.
Magnesium chloride	⅛ lb.
" powdered	⅛ "
" ribbon	⅛ "
" sulphate	¼ "
Manganese chloride	⅛ "
" dioxide	1 "
Marble	5 "
Mercuric chloride	⅛ "
" nitrate	⅛ "
" oxide	¼ "
Mercurous nitrate	⅛ "
Mercury, metallic	½ "
Methyl orange	⅛ oz.
Nickel nitrate	⅛ lb.
Nitric acid	6 "
Oxalic acid	⅛ "
Paraffin	½ "
Phenolphthalein	⅛ oz.
Phosphorus, red	⅛ "
" stick	1 "
Plaster of Paris	⅛ lb.
Potassium bromide	⅛ "
" carbonate	⅛ "
" chlorate	1 "
" chloride	¼ "
" chromate	⅛ "
" cyanide	⅛ "
" dichromate	¼ "
" ferrocyanide	⅛ "
" hydroxide	1 "
" iodide	¼ "
" metallic	⅛ oz.
" nitrate	⅛ lb.
Potassium nitrite	⅛ lb.
" permanganate	⅛ "
" sulphate	⅛ "
" sulphocyanide	¼ "
Silver nitrate	1 oz.
Sodium acetate	⅛ lb.
" carbonate	1 "
" chloride	1 "
" hydroxide	1 "
" metallic	¼ oz.
" nitrate	1 lb.
" nitrite	⅛ "
" phosphate, di-	⅛ "
" sulphate	⅛ "
" sulphide	⅛ "
" sulphite	⅛ "
" thiosulphate	⅛ "
Stannic chloride	⅛ "
Stannous chloride	⅛ "
Starch	¼ "
Strontium chloride	⅛ "
" nitrate	⅛ "
Sugar	½ "
Sulphuric acid	10 "
Sulphur, flowers	1 "
" roll	1 "
Tartaric acid	⅛ "
Tin	⅛ "
" metallic	½ "
Turpentine	⅛ "
Zinc acetate	¼ "
" dust	½ "
" granulated	2 "
" sheet	½ "
" sulphate	½ "

The amounts given above will be found more than sufficient, not only for the students in the laboratory, but also for the lecture room. Except in the case of a few very expensive chemicals, of which very small quantities are needed, it will be economy to buy in larger amounts. The following dealers in chemicals and apparatus are mentioned for the benefit of those making inquiries:

DEALERS IN CHEMICAL SUPPLIES

Eimer & Amend, New York City.
Whitall, Tatum & Co., Philadelphia and Boston.
L. E. Knott Apparatus Co., Boston.
Sargent & Co., Chicago.
Robbins-Martin Co., Chicago.
Henry Heil Co., St. Louis.
Denver Fire Clay Co., Denver.
Bausch & Lomb, Rochester, N.Y.

CHEMICAL APPARATUS

Apron.
Balance and weights (100 g. to 1 dc.).
Beakers (50 and 100 cc.).
Blast lamp.
Blowpipe.
Blue glass.
Bottles, wide mouth (100 cc., 200 cc., 1000 cc.).
Bunsen burners.
Burettes, pinch-cock and stop-cock.
Burette clamps.
" supports.
Clamps, screw and pinch.
Condensers, Liebig, 12 in. and 18 in.
Cork borers, 1 to 6.
Corks, assorted.
Crucibles, porcelain, 1 to $1\frac{1}{2}$ in.
" sand.
Deflagrating spoon.
Electrolytic apparatus.
Eudiometer.
Evaporating dishes, 3 and 4 in.
Files, round and triangular.
Filter pump.
Flasks, 100 cc. and 250 cc.
" " " " (distilling).
Foot bellows.
Funnels, 4 in.
Forceps.
Gauze, iron
Gauze spoon.
Glass rods, 3 mm. diameter.
" spiral.
" tubing, $\frac{1}{4}$ in., $\frac{3}{8}$ in., $\frac{1}{2}$ in. outside diameter.
Graduated cylinder, 200 cc.
Hard glass tubing, $\frac{1}{2}$ in.
Magnet.
Mortar.
Pipette.
Platinum foil, 2 × 2.
" wire.
Pneumatic trough.
Retort, 8 oz.
" stoppered, 12 oz.
Ring-stand and rings.
Rubber corks.
" tubing, $\frac{1}{4}$ in., $\frac{3}{8}$ in.
Sand bath, 3 in.
Separatory funnel.
Test-tubes, 5 and 6 in.
Test-tube brushes.
" clamps.
" supports.
Thermometer, Celsius to 200.
Thistle-tube.
Tripod, iron.
Wire triangle.
Watch-glasses, 4 in.
Water bath.
Wing-top for burner.

WORKING SOLUTIONS

The following table will give in brief the plan of making up solutions for laboratory purposes. For further details see "Modern Chemistry," pp. 376 *et seq.*

To one liter of distilled water, unless stated otherwise, add the following amounts of the substances named: —

Acetic acid, No. 8	1000 cc.
Ammonium carbonate	200 g.
" chloride	100 g.
" hydroxide	3000 cc.
" sulphide	see text.
" molybdate	see text.
Antimony trichloride	10 g. and hydrochloric acid.
Arsenic chloride	see text.
Barium chloride	100 g.
Bismuth nitrate	20 g. and nitric acid to acidulate strongly.
Bleaching powder	saturated solution.
Calcium chloride	100 g.
" sulphate	saturated solution.
Chromium nitrate	50 g.
Cobalt nitrate	50 g.
Cochineal	10 g. treated with 80 per cent alcohol several days.
Copper nitrate	100 g.
" sulphate	100 g.
Hydrochloric acid	500 cc.
Indigo	treated with strong sulphuric acid several days 1 g. to 10 cc., then add 100 cc. water.
Iron chloride, ferric	100 g.
" sulphate	100 g. Make when desired.
Lead acetate	100 g.
" nitrate	100 g.
Lime (hydroxide)	saturated solution.
Litmus	100 g. in hot water.
Magnesium chloride	100 g.
" sulphate	100 g.
Mercuric chloride	saturated solution.
" nitrate	25 g.
Mercurous nitrate	50 g. and 50 cc. nitric acid.

Methyl orange	2 g.		
Nickel nitrate	25 g.		
Phenolphthalein	1 g. in 100 cc. 50 per cent alcohol.		
Potassium bromide	100 g.	Silver nitrate	50 g.
" chromate	50 g.	Sodium carbonate	100 g.
" cyanide	100 g.	" hydroxide	100 g.
" ferrocyanide	75 g.	" nitrite	100 g.
" dichromate	50 g.	" phosphate	100 g.
" iodide	50 g.	" sulphide	40 g.
" hydroxide	100 g.	" sulphite	200 g.
" nitrite	100 g.	" thiosulphate	100 g.
" permanganate	25 g.	Stannic chloride	100 g.
" sulphocyanide	100 g.		
Stannous chloride	75 g. and 50 cc. hydrochloric acid.		
Strontium chloride	100 g.		
Sulphuric acid	500 cc. with care.		
Zinc sulphate	100 g.		

Nessler solution. — Dissolve 17 g. mercuric chloride in 300 cc. pure water, also 35 g. potassium iodide in 100 cc. of water. Pour the mercuric chloride solution into the other until a precipitate begins to remain after stirring. Next add enough of a 20 per cent solution of caustic soda to make the whole amount one liter. Again add mercuric chloride solution until a precipitate remains after stirring. After the precipitate has settled, decant the clear straw-colored solution into a bottle and keep tightly stoppered.

SOLUBILITIES OF GASES IN WATER

One liter of water at zero C. and 760 mm. will absorb the following amounts of the gases named: —

Air	0.02471	liter	Hydrogen	0.0193	liter
Ammonia	1148.0	"	" sulphide	4.73	"
Carbon dioxide	1.7967	"	Marsh gas	0.0545	"
" monoxide	0.03287	"	Nitrogen	0.0203	"
Chlorine (10 degrees C.)	3.0361	"	" monoxide	1.3052	"
Ethylene	0.2562	"	Oxygen	0.0411	"
Hydrochloric acid	503.0	"	Sulphur dioxide	79.789	"

SOLUBILITIES OF SALTS IN WATER

At zero and 100 degrees C., 100 cc. of pure water will dissolve the number of grams of each substance indicated in the table below: —

Compound	0 C.	100 C.	Compound	0 C.	100 C.
Ammonium chloride,	28.4	73.0	Potassium bromide .	53.4	102.0
" hyd. carb.,	12.0	—	" carbonate,	90.0	156.0
" nitrate .	97.0	—	" chlorate .	3.3	56.0
" sulphate,	71.0	97.5	" chloride .	28.0	56.0
Barium chloride . .	31.0	59.0	" dichromate,	5.0	100.0
" hydroxide .	1.6	50.0	" iodide . .	126.0	200.0
Borax crystals . . .	2.83	200.0	" nitrate .	13.3	246.0
Calcium chloride . .	50.0	155.0	" sulphate .	8.36	26.0
" hydroxide .	0.17	0.08	Sodium carbonate, dry . .	7.0	45.5
" sulphate, anhydrous .	0.2	0.17	" chloride . .	35.7	39.7
Copper sulphate, crystals	31.61	203.32	" hydrogen phosphate	2.5	99.0
Lead nitrate . . .	36.5	127.0	" nitrate . .	73.0	180.0
Magnesium sulphate,	26.23	73.57	" nitrite, exceedingly soluble.		
Manganous sulphate, anhydrous . . .	55.0	53.0	" sulphate, crystals . .	12.16	212.47
Mercuric chloride .	5.0	54.0	Zinc sulphate crystals	115.22	653.6
Potassium alum, crystals	3.3	357.5			

ABSORPTION OF GASES BY CHARCOAL*

(Hunter)

One volume of charcoal at zero C. and 760 mm. pressure will absorb the following volumes of the gases, respectively: —

Ammonia	171.7	Hydrogen phosphide . . .	69.1
Carbon dioxide	67.6	Nitrogen	15.2
" monoxide	21.2	" dioxide	86.3
Cyanogen	107.5	" monoxide	70.5
Ethylene	74.7	Oxygen	17.9
Hydrogen	4.4		

* Charcoal from cocoanut shell.

TABLE OF IONIZATION

(Dr. Alexander Smith, Chicago University)

	per cent		per cent
Acetic acid, normal solution,	0.37	Nitric acid, N.	82.0
Ammonium chloride, "	74.0	" " 62 per cent . .	9.6
" hydroxide, normal	0.37	Potassium acetate	64.0
Barium hydroxide normal .	69.0	" chloride, N. . .	75.0
" " N/64 . .	92.0	" hydroxide, N. . .	77.0
Calcium hydroxide, N/64 .	90.0	" nitrate, N. . . .	64.0
" sulphate, N/100. .	63.0	" sulphate, N. . .	53.0
Carbonic acid, N/25 . . .	0.21	Sodium hydroxide, N. . . .	73.0
Di-sodium phosphate, N/32	83.0	Sulphuric acid, N.	51.0
Hydrochloric acid, N. . . .	78.0	" " 95 per cent .	0.7
" " 35 per cent,	13.6	Water, less than one in a million.	
		Zinc sulphate, N.	24.0

TENSION OF AQUEOUS VAPOR

Degrees, C.	Pressure, mm.	Degrees, C.	Pressure, Atmos.
0	4.6	100	1
5	6.534	120.6	2
10	9.165	133.9	3
15	12.699	144.0	4
20	17.391	159.2	6
30	31.548	170.8	8
40	54.906	180.3	10
50	91.982	188.4	12
60	148.791	195.5	14
70	233.093	201.9	16
80	354.28	207.7	18
90	525.45	213.0	20
100	760.0	224.7	25

EQUIVALENTS OF METRIC UNITS

1 millimeter . . .	= 0.03937 in.	1 decimeter . . .	= 3.93708 in.
1 centimeter . . .	= 0.39371 in.	1 meter	= 39.37079 in.

1 cubic centimeter . . = 0.06103 cu. in.
1 liter, or cubic decimeter = 61.02705 cu. in.

1 milligram . . = 0.01543 grains.
1 gram . . . . = 15.43235 grains.
1 kilogram . . = 15432.34880 grains.
1 cubic centimeter = 16.3 minims.

1 centiliter (10 cc.) = 2 dr. 43 min.
1 deciliter (100 cc.) = 3 oz. 3 dr. 2 min.
1 liter (1000 cc.) . = 1 pt. 15 oz. 1 dr. 43 min.

1 pint . . . . = 20 fl. oz.
1 fl. oz. . . . . . . = 8 dr.
1 dr. . . . . = 60 min.

CRITICAL TEMPERATURES AND PRESSURES

The following table is an extract from Hardin's "Liquefaction of Gases."

Gas	Crit. Temp.	Crit. Press.	Boiling Point	Freezing Point
Acetylene	37.5	68.0	——	——
Ammonia	130.0	115.0	− 33.7	− 75.0
Argon	−121.0	50.6	−187.0	−189.6
Carbon dioxide . .	31.0	75.0	− 78.0	− 65.0
" monoxide .	−141.0	36.0	−190.0	−207.0
Chlorine	141.0	83.4	− 36.6	−102.0
Cyanogen	124.0	61.7	− 21.0	− 34.4
Ethylene	10.0	30.0	−102.5	−169.0
Fluorine	−120.0	40.0	−177.0	——
Hydrochloric acid .	51.0	86.0	− 35.0	−116.0
*Hydrogen	−241.0	15.0	−252.0	−260
" sulphide.	100.0	92.0	− 61.8	− 85.0
Marsh gas	− 95.5	50.0	−164.0	——
Nitric oxide . . .	− 93.5	71.2	−153.6	−167.0
Nitrogen	−146.0	35.0	−194.5	−214.0
Nitrous oxide. . .	− 35.4	75.0	− 87.9	−115.0
Oxygen	−118.0	50.0	−183.0	——
Ozone	——	——	−125.0	——
Phosphine	——	——	− 85.0	−133.0
Sulphur dioxide . .	155.4	78.9	− 8.0	——
Water	358.1	——	100.0	0.0

* Dewar, 1899.

LIQUEFACTION OF GASES

Ammonia . . .	At 0 C. and 4.2 atmos. or 15.5 C. under 6.9 pressure.
Carbon dioxide .	−5 C. and 30.8 atmos. or 15 C. under 52.1 pressure.
Chlorine	0 C. and 6.0 atmos. or −34 C. under 1 pressure.
Ethylene . . .	0 C. and 41.0 atmos. or −103 C. under 1 pressure.
Hydrochloric acid	−4 C. and 25 atmos. or 10 C. under 40 pressure.
Nitric oxide . .	−97.5 C. and 57.8 atmos. or −105 C. under 41 pressure.
Nitrous oxide . .	0 C. and 30 atmos.
Nitrogen	−146 C. and 35 atmos. or −148.2 C. under 31 pressure.
Oxygen	−113 C. and 50 atmos. or −129.6 C. under 27 pressure.
Sulphur dioxide .	0 C. and 1.53 atmos. or −10 C. under 1 pressure.

NAMES AND FORMULAS OF THE MORE COMMON CHEMICALS

Acetic acid	$HC_2H_3O_2$
Alcohol	C_2H_5OH
Alum, ammonium	$(NH_4)_2Al_2(SO_4)_4$
" , chrome	$K_2Cr_2(SO_4)_4$
" , potassium	$K_2Al_2(SO_4)_4$
Aluminum oxide	Al_2O_3
" sulphate	$Al_2(SO_4)_3$
Ammonium bicarbonate	NH_4HCO_3
" carbonate	$(NH_4)_2CO_3$
" chloride	NH_4Cl
" hydroxide	NH_4OH
" molybdate	$(NH_4)_2MoO_4$
" nitrate	NH_4NO_3
" oxalate	$(NH_4)_2C_2O_4$
" sulphate	$(NH_4)_2SO_4$
Antimony oxychloride	$SbOCl$
" sesquioxide	Sb_2O_3
" trichloride	$SbCl_3$
" trisulphide	Sb_2S_3
Arsenic trichloride	$AsCl_3$
" trioxide	As_2O_3
" trisulphide	As_2S_3
Barium carbonate	$BaCO_3$
" chloride	$BaCl_2$
" chromate	$BaCrO_4$
" dioxide	BaO_2

Barium hydroxide	$Ba(OH)_2$
" nitrate	$Ba(NO_3)_2$
" oxide	BaO
" sulphate	$BaSO_4$
Bismuth chloride	$BiCl_3$
" nitrate	$Bi(NO_3)_3$
" oxychloride	$BiOCl$
" subnitrate	$BiONO_3$
" trioxide	Bi_2O_3
Bleaching powder	$Ca(ClO)_2 + CaCl_2$(?)
Borax	$Na_2B_4O_7$
Cadmium nitrate	$Cd(NO_3)_2$
" sulphide	CdS
Calcium carbide	CaC_2
" carbonate	$CaCO_3$
" chloride	$CaCl_2$
" fluoride	CaF_2
" hydroxide	$Ca(OH)_2$
" oxide (lime)	CaO
" phosphate	$Ca_3(PO_4)_2$
" sulphate	$CaSO_4$
Carbolic acid	C_6H_5OH
Carbon disulphide	CS_2
Chloroform	$CHCl_3$
Chrome yellow	$PbCrO_4$
Chromium nitrate	$Cr(NO_3)_3$
" sulphate	$Cr_2(SO_4)_3$
" trixoide	CrO_3
Cinnabar	HgS
Cobalt chloride	$CoCl_2$
" nitrate	$Co(NO_3)_2$
Copper acetate	$Cu(C_2H_3O_2)_2$
" chloride	$CuCl_2$
" nitrate	$Cu(NO_3)_2$
" oxide	CuO
" sulphate	$CuSO_4$
" sulphide	CuS
Ether, sulphuric	$(C_2H_5)_2O$
Ferric chloride	Fe_2Cl_6
" hydroxide	$Fe_2(OH)_6$
" nitrate	$Fe_2(NO_3)_6$

Ferric oxide	Fe_2O_3
Ferrous sulphate	$FeSO_4$
" sulphide	FeS
Fluor spar	CaF_2
Gold trichloride	$AuCl_3$
Hydrochloric acid	HCl
Hydrofluoric acid	HF
Hydrogen peroxide	H_2O_2
" sulphide	H_2S
Hypochlorous acid	$HClO$
Iodic acid	HIO_3
Iodoform	CHI_3
Lead acetate	$Pb(C_2H_3O_2)_2$
" carbonate	$PbCO_3$
" chloride	$PbCl_2$
" chromate	$PbCrO_4$
" nitrate	$Pb(NO_3)_2$
" oxide (litharge)	PbO
" sulphate	$PbSO_4$
Lime	CaO
Litharge	PbO
Lithium chloride	$LiCl$
Magnesia	MgO
Magnesium carbonate	$MgCO_3$
" chloride	$MgCl_2$
" oxide	MgO
" sulphate	$MgSO_4$
Manganese chloride	$MnCl_2$
" dioxide	MnO_2
" hydroxide	$Mn(OH)_2$
" sulphate	$MnSO_4$
Mercuric chloride	$HgCl_2$
" iodide	HgI_2
" nitrate	$Hg(NO_3)_2$
" oxide	HgO
" sulphate	$HgSO_4$
" sulphide	HgS
Mercurous chloride	Hg_2Cl_2
" iodide	Hg_2I_2
" nitrate	$Hg_2(NO_3)_2$
Minium	Pb_3O_4

Nickel chloride	$NiCl_2$
" hydroxide	$Ni(OH)_2$
" nitrate	$Ni(NO_3)_2$
" sulphide	NiS
Phosphine	PH_3
Phosphorus pentoxide	P_2O_5
" trichloride	P_2O_3
Plaster of Paris	$CaSO_4$, H_2O
Platinum tetrachloride	$PtCl_4$
Potassium acetate	$KC_2H_3O_2$
" bicarbonate	$KHCO_3$
" bromide	KBr
" carbonate	K_2CO_3
" chlorate	$KClO_3$
" chloride	KCl
" chromate	K_2CrO_4
" cyanide	KCy
" dichromate	$K_2Cr_2O_7$
" ferricyanide	K_3FeCy_6
" ferrocyanide	K_4FeCy_6
" hydroxide	KOH
" iodide	KI
" nitrate	KNO_3
" nitrite	KNO_2
" permanganate	$KMnO_4$
" silicate	K_4SiO_4
" sulphate	K_2SO_4
" sulphocyanide	$KSCy$
Silica	SiO_2
Silver bromide	$AgBr$
" chloride	$AgCl$
" iodide	AgI
" nitrate	$AgNO_3$
" nitrite	$AgNO_2$
Sodium acetate	$NaC_2H_3O_2$
" arseniate	Na_3AsO_4
" arsenite	Na_3AsO_3
" bicarbonate	$NaHCO_3$
" carbonate	Na_2CO_3
" chloride	$NaCl$
" hydroxide	$NaOH$

Sodium iodide	NaI
" nitrate	$NaNO_3$
" nitrite	$NaNO_2$
" phosphate	Na_3PO_4
" sulphate	Na_2SO_4
" sulphide	Na_2S
" sulphite	Na_2SO_3
" thiosulphate	$Na_2S_2O_3$
Stannic chloride	$SnCl_4$
" oxide	SnO_2
Stannous chloride	$SnCl_2$
Strontium carbonate	$SrCO_3$
" chloride	$SrCl_2$
" hydroxide	$Sr(OH)_2$
" nitrate	$Sr(NO_3)_2$
" oxide	SrO
Sulphur dioxide	SO_2
Sulphuric acid	H_2SO_4
Sulphurous acid	H_2SO_3
Sulphur trioxide	SO_3
Zinc carbonate	$ZnCO_3$
" chloride	$ZnCl_2$
" oxide	ZnO
" sulphate	$ZnSO_4$
" sulphide	ZnS

TABLE OF THE ELEMENTS AND THEIR ATOMIC WEIGHTS

Name	Symbol	Atomic Weights	
		O = 16	H = 1
Aluminum	Al	27.1	26.9
Antimony	Sb	120.	119.5
Argon	A	39.9	?
Arsenic	As	75.	74.45
Barium	Ba	137.4	136.4
Bismuth	Bi	208.5	206.5
Boron	B	11.	10.9
Bromine	Br	79.96	79.34
Cadmium	Cd	112.4	111.55
Cæsium	Cs	133.	131.9
Calcium	Ca	40.	39.8
Carbon	C	12.	11.9
Cerium	Ce	140.	138.0
Chlorine	Cl	35.45	35.18
Chromium	Cr	52.1	51.7
Cobalt	Co	59.	58.55
Columbium	Cb	94.	93.0
Copper	Cu	63.6	63.1
Erbium	E	166.	164.7
Fluorine	F	19.	18.9
Gadolinium	Gd	156.	155.8
Gallium	Ga	70.	69.5
Germanium	Ge	72.	71.9
Glucinum	Gl	9.1	9.0
Gold	Au	197.2	195.7
Helium	He	4.	?
Hydrogen	H	1.01	1.0
Indium	In	114.	113.
Iodine	I	126.85	125.89
Iridium	Ir	193.	191.7
Iron	Fe	56.	55.5
Krypton	Kr	81.8	?

TABLE OF THE ELEMENTS AND THEIR ATOMIC WEIGHTS — *Continued*

Name	Symbol	Atomic Weights	
		O = 16	H = 1
Lanthanum	La	138.	137.6
Lead	Pb	206.9	205.36
Lithium	Li	7.	6.97
Magnesium	Mg	24.36	24.1
Manganese	Mn	55.	54.6
Mercury	Hg	200.3	198.50
Molybdenum	Mo	96.	95.3
Neodymium	Nd	143.6	142.5
Neon	Ne	20.	?
Nickel	Ni	58.7	58.25
Nitrogen	N	14.04	13.93
Osmium	Os	191.	189.6
Oxygen	O	16.	15.88
Palladium	Pd	106.	106.2
Phosphorus	P	31.	30.75
Platinum	Pt	194.8	193.4
Potassium	K	39.15	38.82
Praseodymium	Pr	140.5	139.4
Rhodium	Rh	103.	102.2
Rubidium	Rb	85.4	84.75
Ruthenium	Ru	101.7	100.9
Samarium	Sm	150.	149.2
Scandium	Sc	44.1	43.8
Selenium	Se	79.1	78.6
Silicon	Si	28.4	28.2
Silver	Ag	107.93	107.11
Sodium	Na	23.05	22.88
Strontium	Sr	87.6	86.95
Sulphur	S	32.06	31.83
Tantalum	Ta	183.	181.5
Tellurium	Te	127.	126.5
Terbium	Tr	160.	158.8

TABLE OF THE ELEMENTS AND THEIR ATOMIC WEIGHTS

Name	Symbol	Atomic Weights	
		O = 16	H = 1
Aluminum	Al	27.1	26.9
Antimony	Sb	120.	119.5
Argon	A	39.9	?
Arsenic	As	75.	74.45
Barium	Ba	137.4	136.4
Bismuth	Bi	208.5	206.5
Boron	B	11.	10.9
Bromine	Br	79.96	79.34
Cadmium	Cd	112.4	111.55
Cæsium	Cs	133.	131.9
Calcium	Ca	40.	39.8
Carbon	C	12.	11.9
Cerium	Ce	140.	138.0
Chlorine	Cl	35.45	35.18
Chromium	Cr	52.1	51.7
Cobalt	Co	59.	58.55
Columbium	Cb	94.	93.0
Copper	Cu	63.6	63.1
Erbium	E	166.	164.7
Fluorine	F	19.	18.9
Gadolinium	Gd	156.	155.8
Gallium	Ga	70.	69.5
Germanium	Ge	72.	71.9
Glucinum	Gl	9.1	9.0
Gold	Au	197.2	195.7
Helium	He	4.	?
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Selenium	Se	79.1	78.6
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Silver	Ag	107.93	107.11
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Terbium	Tr	160.	158.8

TABLE OF THE ELEMENTS AND THEIR ATOMIC WEIGHTS — *Continued*

Name	Symbol	Atomic Weights	
		O = 16	H = 1
Thallium	Tl	204.1	202.61
Thorium	Th	232.5	230.8
Thulium	Tm	171.	169.4
Tin	Sn	118.5	118.1
Titanium	Ti	48.1	47.8
Tungsten	W	184.	182.6
Uranium	U	239.5	237.8
Vanadium	V	51.2	51.0
Xenon	X	128.	?
Ytterbium	Yb	173.	171.9
Yttrium	Y	89.	88.3
Zinc	Zn	65.4	64.9
Zirconium	Zr	90.7	89.7

The above table shows two columns of atomic weights; the first assumes O = 16 as the standard, the second, H = 1.

PERIODIC TABLE OF ELEMENTS

		I	II	III	IV	V	VI	VII	VIII
Period I		H=1 Li=7	Gl=9	B=11	C=12	N=14	O=16	F=19	
" II		Na	Mg	Al	Si	P	S	Cl	
" III	A	K	Ca	Sc	Ti	V	Cr	Mn	Fe, Co,
	B	Cu	Zn	Ga	Ge	As	Se	Br	Ni
" IV	A	Rb	Sr	Y	Zr	Cb	Mo		Ru, Rh,
	B	Ag	Cd	In	Sn	Sb	Te	I	Pd
" V	A	Cs	Ba	La	Ce				
	B								
" VI	A			Yb			W		Os, Ir,
	B	Au	Hg	Tl	Pb	Bi			Pt
" VII	A			Th			U		
	B								

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